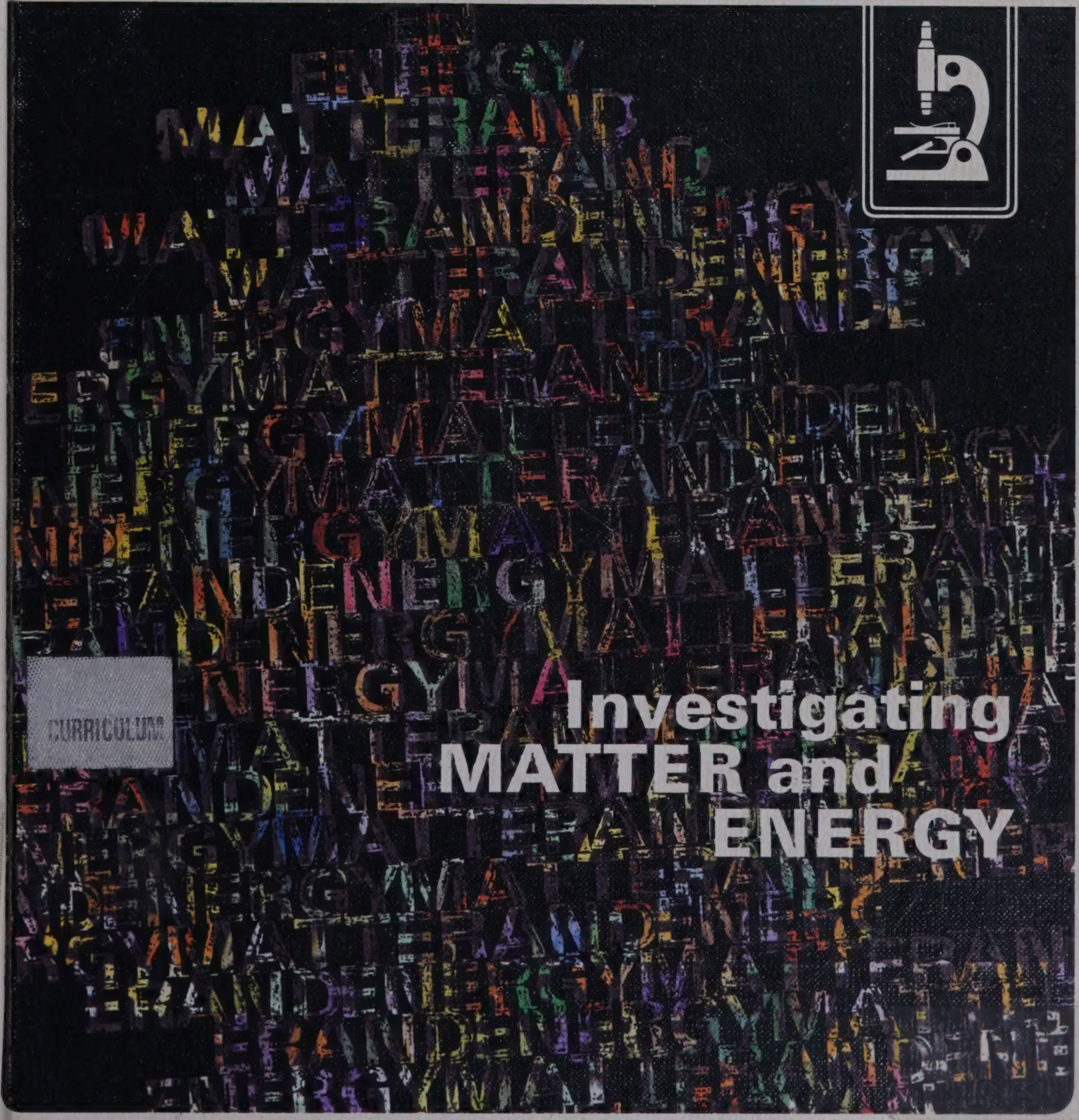
PHYSICAL SCIENCE

A LABORATORY APPROACH



Investigating MATTER and ENERGY

CURRICULUM

PHYSICAL SCIENCE
A LABORATORY APPROACH

Investigating
MATTER and
ENERGY

John H. Marean
Elaine W. Ledbetter

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Investigating MATTER and ENERGY

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Preface

The recent revolution in science education has been stimulated by the explosive advance of twentieth century science. The initial leaders of this revolution were scientists. Their efforts have led to profound changes in the content of high school science courses. The added efforts of science teachers, in cooperation with scientists, have brought about marked changes in the methods of science teaching.

The influence of this new outlook in science education has been most evident in senior high schools. From there it has spread to other levels. College programs have been reorganized to utilize the new attitudes and learnings acquired by students. Teachers of elementary school and junior high school science are attempting to prepare their students for the new method and content that they will encounter in high school science.

Beginning in the fall of 1962 we began to develop a new approach to teaching junior high school students the basic concepts of physical science. We felt that the student should be acquainted with science as an *activity* rather than as a *product*. In physical science variables are easily controlled and actual experimentation is possible with very simple equipment. The consequences of the students' explorations can readily lead to interpretation of concepts that are worthwhile as a general background as well as fundamental to further study of science.

The goal of this entire course is to provide a plan by which the student may systematically, through relatively simple and meaningful investigation, build a conceptual understanding of the structure of matter and the nature of energy. The student is confronted with critical questions as he builds his concept of these two big ideas. The suggested activities have been chosen because they will help build confidence in one or another theory of the structure of matter or of the nature of energy. In the long run the student develops a good grasp of the credibility of the accepted theories of matter and energy and of their interrelationships in a number of basic areas of study.

In such a pioneer venture as this it was extremely important to make the best use of the experience and thoughtful suggestions of both teachers and scientists. We want to thank those whose talents and energies have helped us to make this program generally useful to teachers. Miss Helen Hale and Mr. Samuel Herman of Baltimore County, Maryland, and Mr. Richard Schultz of Cedar Rapids, Iowa, as administrators, provided opportunities for field-testing this program. Mr. Douglas Branch and Mrs. Willie Mae Mangold, both of Pampa, Texas, Mr. Dudley Cate and Mr. Edward Coppin, both of Reno, Nevada, and Mr. Donald Patrick of Baltimore County, Maryland, used and evaluated the preliminary editions of this program in their classrooms. Numerous others have provided helpful suggestions.

J.H.M.
E.W.L.

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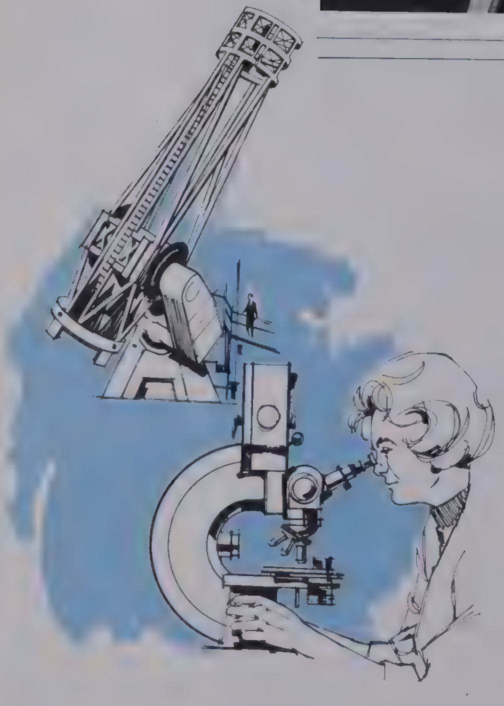
1

The Nature of Science



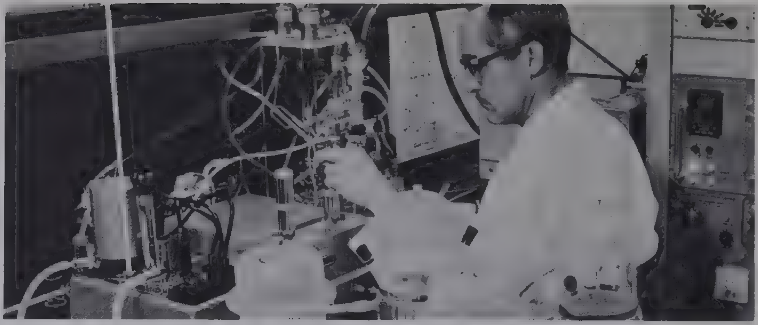
1–1 Science and Curiosity
1–2 The Process and Results
1–3 Drawing Conclusions From Indirect Evidence
1–4 The Tools of Science
1–5 The Metric System
1–6 Induction and Deduction

You are a reliable source of information.

1–1 SCIENCE AND CURIOSITY

As a fundamental part of his makeup, man is curious. Think how many times you have asked your parents, teachers, or friends a question to satisfy your curiosity about "what" or "how" or "why." A scientist is a person who has made curiosity his profession. He asks questions in many ways, gets answers from many sources, and uses all existing data and past experience to judge the merit of the answers. In this respect you have solved problems in the same manner a scientist would. You have asked questions in many ways and received answers from many sources. You have judged the merit of the answers you got by your opinion of the source. Your teacher, as one source, may be able to give you much valuable information and help you with your schoolwork and still be unable to tell you what to expect for dinner tonight. Your mother, as another source, can probably answer your questions about dinner, but may not be able to help with your homework. You, too, are a source of information. Your senses of smell and sight may disclose the menu, and your memory and intelligence may provide the needed information for school. Each source is reliable for one kind of answer and each one should be used in its proper place.

The scientist who devotes his life to seeking the most meaningful answers to the questions that fire his curiosity strives constantly for the best sources of information. His curiosity is highly directed. He guides it, puts controls on it, and subjects each of the answers he gets to the most careful tests to be sure it is the best answer he can find. For this reason we have confidence in scientific information.



Both scientists and carpenters need special equipment and special places for much of their work.

1–2 THE PROCESS AND RESULTS

Science is twofold in nature: It is the *process* of asking and trying to answer the most meaningful questions, and it is the *results* of this process—the facts of science.

Science is the work of scientists, just as carpentry is the work of carpenters. We can look at results of both kinds of work. But we cannot learn much about the methods and techniques of either a scientist or a carpenter by looking only at their finished products. Their tools and their techniques must be examined directly.

We could learn about science by watching a scientist at work. But it is not easy to follow a scientist through his tasks. The techniques of science are as numerous as the problems to be solved, and all too often we learn more about the results achieved than the techniques used. Thus, in order to understand the activity called science, we cannot merely observe scientists at work. We must ourselves participate in scientific experiments.

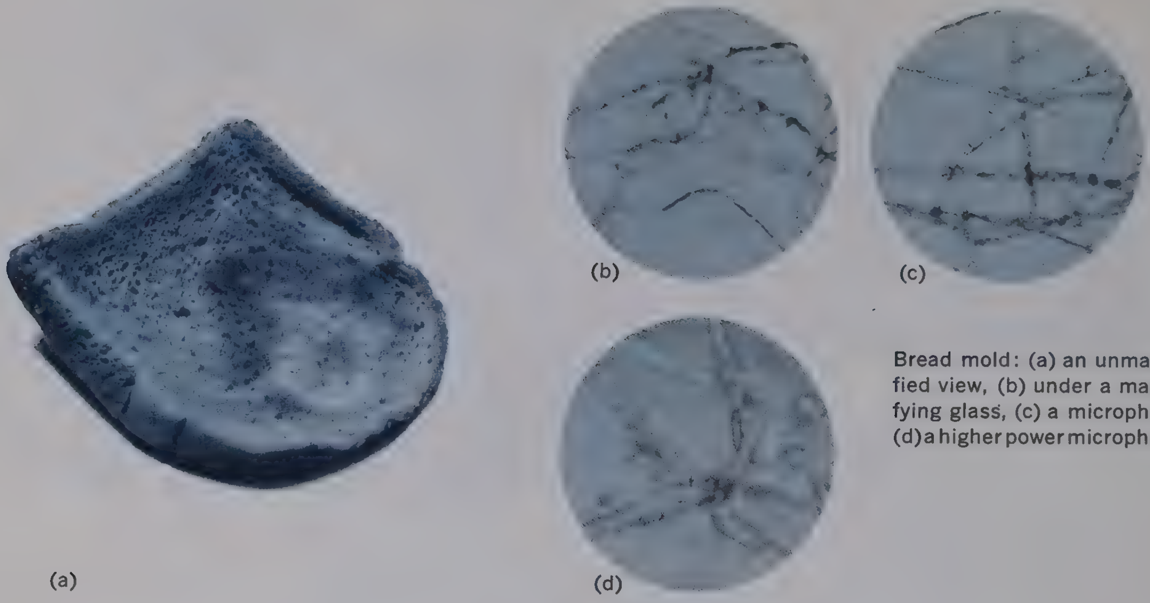
An experiment begins in the mind of a scientist. He wants to know about something that has aroused his interest and curiosity. He has a *question* that he wants answered. Along with the question, he may have some ideas about answers. Sometimes these answers are wild guesses, but more often the scientist has good reason for thinking that one of the answers is correct. The scientist's best guess as to the outcome of an experiment, the answer to his question, is called a *hypothesis*. To decide

whether his hypothesis is valid, he must devise an experiment—a carefully designed *attack* on the question. If the attack is well planned, the experiment will produce worthwhile information that the scientist can use in supporting or discarding his hypothesis.

A Controlled Experiment

To illustrate a controlled experiment and to gain experience in making direct observations, let's ask a question and proceed to answer it as a scientist might. There is no single experiment that would answer the question, What causes bread to mold? To learn the answer we have to ask many questions. Since bread often molds in the conditions that exist where we store it in our homes, we need to consider each of these conditions separately, and from the answers we get to individual questions, decide what conditions seem to contribute to the growth of mold in bread. Such questions might be: Does the temperature affect the growth of mold on bread? Does the amount of moisture present affect the growth of mold on bread? Does the amount of light present affect the growth of mold on bread? Does the growth of mold on bread depend on time? Does bread mold develop and grow as readily on the crust as on the face of a slice of bread? Do homemade baked goods seem to mold as readily as those produced by commercial bakeries?

In each of these questions, there are only two possible answers. With a carefully designed experiment we can probably conclude whether



Bread mold: (a) an unmagnified view, (b) under a magnifying glass, (c) a microphoto, (d) a higher power microphoto.

any one question may be answered yes or no. But we cannot, with one experiment, answer all six. In addition, we need to word our questions in such a way that we believe we know what the answer should be. Let's choose one question from the list given in the preceding paragraph.

Question: Does the amount of moisture present affect the growth of mold on bread? (The skill we exercise in presenting the question greatly affects the quality of the answer we receive.)

Hypothesis (your best guess): Whether you choose yes or no as your best guess should depend on your previous experience. 1 Can you recall, as you have observed bread mold in the past, the conditions present that might have been influences?

Attack: Here is a suggestion of procedure to follow to produce some reliable results which will help us conclude whether the answer is yes or no. To carry out this procedure you will need two slices of bread from the same loaf, two widemouth jars with tight-fitting lids, two labels, and some water.

1. Moisten one slice of bread with water and put it in a jar. Screw on the lid to keep the bread from drying out. Label this jar "moist bread" and write the date on the label.

2. Put the second slice of bread in the other jar. Screw on the lid and label the jar "unmoistened bread." Write the date on the label.

3. Place the jars where they will not be disturbed. They should be side by side so that the other conditions, such as light and temperature, will be the same.

4. Examine the bread in both jars each day and keep a record of your observations.

Results: $\boxed{2}$ On the basis of what you have observed, does it appear that moisture affects the growth of mold on bread?

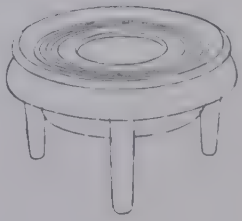
One important part of research in science is to be sure we can make accurate comparisons when we have finished making observations. To do this we must establish controls in the experiment. Whenever possible you should prepare two setups as nearly alike as possible in every way with the exception of one factor. You just did this in Experiment 1–1 by putting two slices of bread from the same loaf in two similar jars with lids, and by placing both jars in the same environment. The only variable factor was that one slice of bread was moistened with water, while the other slice was not moistened. The unmoistened bread is known as the *control.* This use of a control makes it possible to be more certain that any differences observed are caused by the variable factor.

Our entire experiment depended on the nature of our hypothesis—on the guess we made. We guessed that moisture had something to do with the growth of bread mold. We need not experiment to know that there are some hypotheses that are not valid. The growth of mold obviously does not depend directly on the price of bread, or on the height of the Rocky Mountains. $\boxed{3}$ Does it depend, in any direct way, on the season of the year or on the phase of the moon? No guess should be discarded without thought, but neither should a guess be accepted as true because it seemed true as we thought about it. If we do not have experience with the situation, the question must be subjected to experimentation before we can draw a reliable conclusion.

In drawing conclusions we must be careful to base our statements on the actual results observed in the experiment. $\boxed{4}$ How many times should an experiment like this be repeated before it is safe to make a general statement about the effect of moisture on the growth of mold on bread? $\boxed{5}$ How do you know that it was moisture itself and not an ingredient of the water added that produced any difference? $\boxed{6}$ How would you find out? [A, B, C]

Keeping Good Records

The data from each experiment should be recorded in a manner that is easy to read and understand. All the available information that may help solve the problem should be listed in such a way that it will be easy to use. A carpenter, for example, keeps various kinds of lumber in separate stacks and does not store all different kinds of nails in the same box. He keeps the materials for his project near one another to avoid



unnecessary trouble in organizing them and to avoid overlooking something. He does not clutter the job with scrap.

The scientist works in a similar way. When he completes an experiment he assembles his evidence into clear, understandable patterns. He uses words, numbers, drawings, and diagrams to communicate his findings. In finished form, his report is published so that other scientists will know of his work.

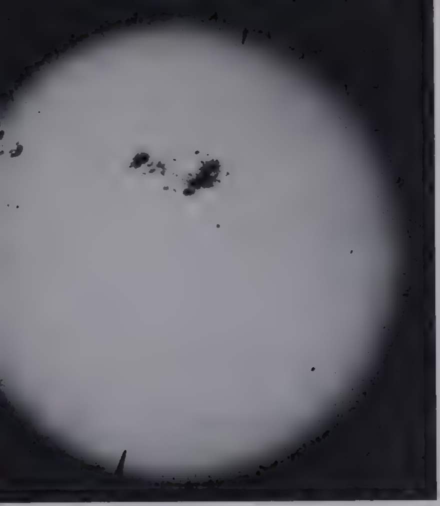
Throughout this course you will be conducting experiments and keeping records of them. One of the most satisfactory ways to organize your material is to keep a laboratory record book. It is better to record too much information than too little. You should date and enter every observation *at the time it is made.* You may have ideas and suggestions that you want to record as well as actual observations. It will be necessary to note how your experiments are related to the main idea of a chapter or a section so that you do not lose sight of why you are making certain observations.

You will find that as you gain experience in laboratory work, you will also gain skill in recording observations. Sometimes making a chart will be the most useful way to organize your data. In other investigations, a list of results may be more meaningful. Your teacher can suggest a preferred method of recording data for a specific observation if you are unsure of how to do it. Whatever the method, make sure that you date your records of each activity. In later chapters you will be asked to refer to data collected early in the course; the dates will help you find the information as you need it.

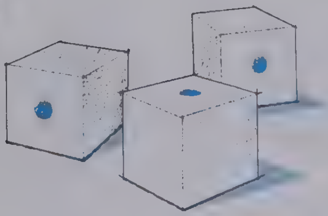
1–3 DRAWING CONCLUSIONS FROM INDIRECT EVIDENCE

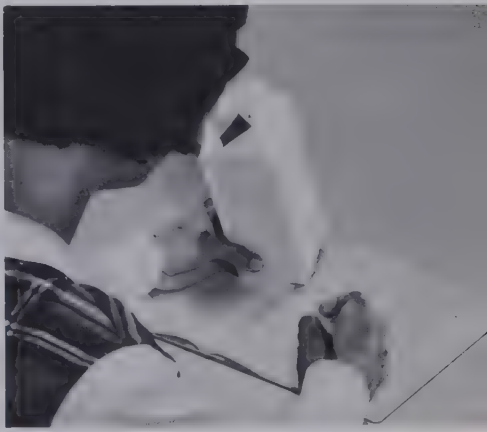
Direct evidence such as you obtained from the bread mold experiment is rather easy to get. But very often scientists must obtain evidence indirectly. We know much about the shape, size, composition, temperature, and rotation of the sun. Yet no one has been to the sun, and it is unlikely that man will ever be able to go there. Our information about the sun has been acquired by *indirect methods.*

If we pose a simple problem and use indirect methods in solving it, we may learn more about the activities of science. Experience in gathering evidence indirectly may help us to use similar work methods to solve other problems. Imagine that you are going to learn something special about markings on a sugar cube without touching it and only seeing one surface at a time. Only the top of the cube shows in its rest position in the bottom of a tall, narrow container, and as you look at it, you see a spot



Our knowledge of sunspots is based on indirect evidence.





What can you predict about possible spots on the faces of each cube?

on this surface. You might wonder if there is a spot on any of the other five faces. Since you cannot touch the cube to turn it, you must shake the container and try to shift the position of the cube to show another face. To make the situation more complex, but perhaps more like that which a scientist actually meets, let's add the condition that you may not observe the cube during the shaking process.

Question: How many faces on the cube are marked?

Hypothesis: The number of faces marked cannot be less than zero nor more than six and it must be an integral (whole) number.

Attack: One way to gather information about a cube is to use a procedure such as the following. You will need a tall container (such as a clean dry milk carton with the top removed), a sugar cube marked with a spot on a number of faces unknown to you, and your laboratory record book.

1. Put the sugar cube into the tall, narrow container without examining it.

2. Look at the top face of the cube as it rests in the container to determine if it is marked with a spot. Record your observation.

3. Shake the container. After the cube has come to rest, examine it again and record your observation.

4. Repeat this procedure until 100 observations have been made and recorded.

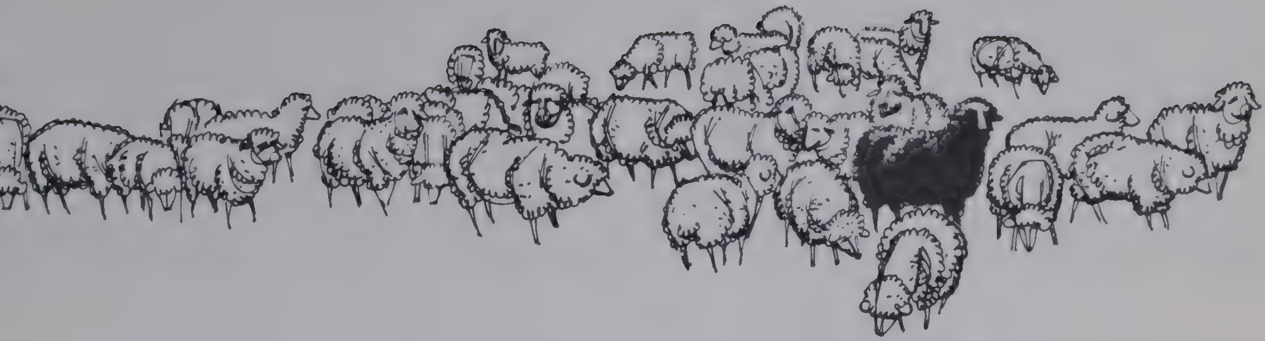
TABLE 1-1

22 September

Indirect evidence

1. Received open-top milk carton containing sugar cube.
2. Observed top face of cube.
3. Recorded observation of marking.
4. Shook carton to change cube and repeated observation and recording.
5. Repeated for a total of 100 readings.

Trial number	Spot observed	No spot observed	Trial number	Spot observed	No spot observed	Trial number	Spot observed	No spot observed	Trial number	Spot observed	No spot observed
1		✓	26	✓		51		✓	76		✓
2		✓	27		✓	52		✓	77	✓	
3		✓	28		✓	53		✓	78	✓	
4	✓		29		✓	54		✓	79	✓	
5		✓	30	✓		55		✓	80	✓	
6	✓		31		✓	56	✓		81		✓
7	✓		32	✓		57		✓	82		✓
8		✓	33		✓	58		✓	83		✓
9		✓	34	✓		59	✓		84		✓
10		✓	35		✓	60		✓	85	✓	
11		✓	36		✓	61	✓		86		✓
12	✓		37	✓		62	✓		87	✓	
13	✓		38		✓	63	✓		88		✓
14		✓	39		✓	64		✓	89	✓	
15		✓	40	✓		65	✓		90	✓	
16		✓	41	✓		66		✓	91		✓
17		✓	42		✓	67		✓	92		✓
18		✓	43		✓	68	✓		93		✓
19		✓	44		✓	69	✓		94		✓
20		✓	45		✓	70		✓	95		✓
21		✓	46		✓	71	✓		96	✓	
22	✓		47	✓		72		✓	97		✓
23		✓	48		✓	73	✓		98	✓	
24		✓	49		✓	74		✓	99		✓
25		✓	50		✓	75		✓	100		✓



Results: The observations recorded in Table 1–1 came from the study of such a sugar cube. These are given in order that you may see how to record your own data and learn how to draw conclusions from your data. **Analysis:** During the 100 trials we looked at 100 faces. Counting the total number of times a spot was observed, we find there were 34. When we divide the total number of spots observed (34) by the total number of faces observed (100), we have 0.34 spots per face. The cube has 6 faces, so to find the probable number of spots on the whole cube we multiply the number of spots per face (0.34) by the number of faces (6) and find there are 2.04 spots on the cube. We know this is an *average* and, since either a whole spot or no spot was seen, we drop the decimal and conclude that there were probably two faces that had a spot and four faces that were not marked.

How much evidence is enough? Look at the results of the first six trials and you will notice that a spot appeared only twice. $\boxed{7}$ Would it be safe to stop making observations after just six trials? $\boxed{8}$ What if the first six trials had given the results that were observed in trials 61–66? On the basis of these six trials, we would conclude that there were four faces marked with a spot and two that were unmarked. $\boxed{9}$ What conclusion would you draw if you looked only at trials 15–20?

Obviously, when we are trying to draw conclusions from this kind of observation, we must include results of enough tests so that these small variations are lost in the large average. We must be very careful to consider enough cases so that a reliable average exists. After 100 trials you should have confidence in your ability to describe how many spots are on a cube. $\boxed{10}$ How many trials are required for complete confidence?

The importance of caution in interpreting data can be seen in the following example. For centuries shepherds recognized that, on the average, one out of every 100 lambs born is black. This "rule of thumb" has been used by shepherds to give the approximate number of sheep in a flock. But assume that, during a night raid by a pack of wolves on a large flock of sheep, two of the six sheep killed happened to be black. [11] Should the shepherd decide that wolves had killed 200 of his sheep?

Conclusions: How many faces of the cube reported in Table 1–1 really had a spot? We shall never know for certain, because it fell from the container into the author's coffee before there was an opportunity to examine it in detail. [D, E, F]

A Study of Averages

1 3 Scientists seldom determine properties by examining a single object. More often they examine a great number and find the average for the group. By experimenting carefully and recording a number of readings, we may learn something about a group of objects. Let's use sugar cubes again and shake a group of them together. Put four or five cubes together in one container. This time, record the *number of cubes* on which a spot is observed after each shaking. Here is one test with five cubes in the container. The following observations were made:

A spot was observed on each of the five cubes 3 times;	$5 \times 3 = 15$
A spot was observed on four of the five cubes 11 times;	$4 \times 11 = 44$
A spot was observed on three of the five cubes 24 times;	$3 \times 24 = 72$
A spot was observed on two of the five cubes 42 times;	$2 \times 42 = 84$
A spot was observed on one of the five cubes 14 times;	$1 \times 14 = 14$
No spots were observed on any of the cubes 6 times;	$0 \times 6 = 0$
	229

[12] What was the average number of spots on each of the cubes? This number does not have to be an integer, even though the actual number on each of the cubes is an integer. Let's analyze our findings. In all, with five cubes and 100 observations we saw a total of 500 faces ($5 \times 100 = 500$). Of these, 229 carried a spot. This means that 45.8 percent of the faces we observed had spots ($\frac{229}{500} = 0.458$). At six faces per cube, the average is 2.75 spots per cube (0.458×6). This does not tell us anything conclusive about any *one* of the cubes. We may say quite confidently that none of the cubes had a spot on each face and also that none of the cubes was entirely unmarked. [13] Why can we say this? [14] Is it possible that each of the cubes had three spots? [15] Are there other marking combinations that are equally probable? If we also had the same kind of data we gained earlier on one or more of these five, we could make a more accurate statement about the remaining cubes. [G]

Two examples of modern classification.

1–4 THE TOOLS OF SCIENCE

Learning all we can about all that we observe is surely an important phase of science and, consequently, a task of the scientist. Think, though, of the vast amount of information we could assemble and of how difficult it could become to find a certain item of this information if we did not group it in some convenient manner. Consider how much simpler our study will be if we combine the objects we are studying into classes or kinds before we begin to gather detailed data. By doing this we may generalize about many members of the group after a careful study of only a few members. In effect, we did this in our experiment on the single sugar cube when we grouped our observations. We said that in 100 trials a spot appeared in 34 and no spot in 66. This classification made it simpler to describe what we had observed and easier for us to evaluate our data in order to draw conclusions about what could not be observed directly. This grouping, or sorting, activity is the science of classification; it is called *taxonomy*.

Classification is a sorting operation and can be done according to many different systems. The plan we use depends on our purpose. The extent of sorting also varies with the situation. Your grocer keeps certain kinds of food together in the store for your shopping convenience and for his convenience in restocking supplies. The cans of food you keep at home may be sorted in the cabinet with fruits, vegetables, and soups on different shelves, or at different places on a single shelf. A large family storing many cans of food may have a more detailed system of classification than a small family with fewer cans.

The library offers an excellent example of how we may sort and classify objects. Here, a large variety of reading material is grouped into books, magazines, and newspapers. The books are further divided into subgroups such as fiction and nonfiction. Nonfiction is broken down into more specific, smaller groups such as books on science, history, and hobbies. Such organized sorting makes for more efficient and convenient use of the materials available.

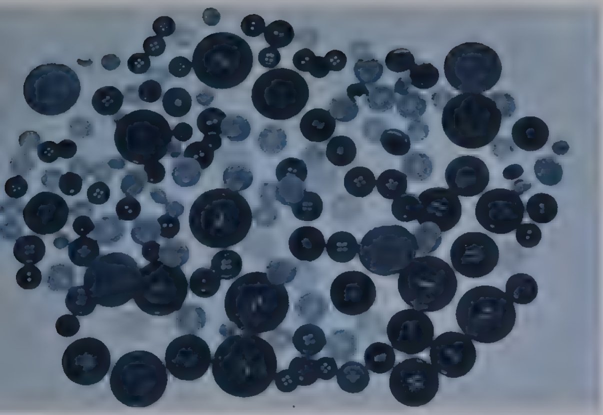
The first problem in a program of classification is to determine the distinctive common characteristics of the members of each subgroup. This decision is made by the taxonomist. Any system is potentially as good as any other, provided that there is reason for the system selected. You are accustomed to working with classification and are already trained in certain systems. Let's look at a few examples of existing systems applied to familiar objects.

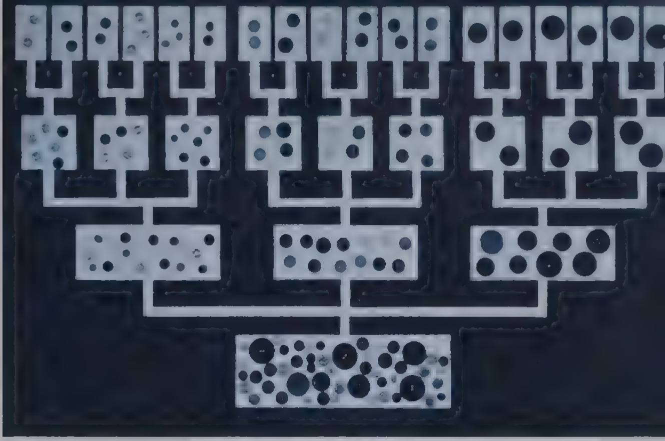
You are well acquainted with transportation devices. All devices used for any kind of transportation can be listed under this general title. Within the large group are vehicles used for transportation on land, in air, on water, under water, and through space. Within each of these divisions many types of devices can be identified. Ground transportation may be divided into groups powered by internal combustion engines, by steam, by electricity, or by animals. They may also be divided according to the number of wheels. Another method of grouping these vehicles depends on whether they carry human passengers, animals, or inanimate cargo. Obviously, our basis of classification determines whether we group certain vehicles together. A bicycle and a train are in the same group if we are concerned with transporting humans, but in different classes if we rank them according to the type of power or the number of wheels or the type of roadbed used. Which system is better? That depends on the purpose for which it is intended.

Each of the examples we have discussed requires a different kind and degree of classification. Note, however, that classification can be done only after we have learned a great deal about the items to be classified. We must know enough about the individual objects to recognize similarities and differences, so that grouping can be done for the greatest convenience. Sometimes we do not physically group our objects or ideas but rather we organize useful knowledge of them. All people by the name of Smith are not actually grouped together, even though their names are grouped together in the telephone directory.

Practice in Classification

It would be well for you to gain practice in classification and to experience the decision-making that must accompany any system of grouping. Buttons are excellent items for practice in classification. Their great





A classification tree for buttons.

variety in kind and color makes it possible to try many different systems of grouping. For the present we may say that no system is better than any other. Until we know the reason for classifying something, any consistent system is good.

Take as many buttons as you can get and sort them into groups so that all the members of any group have some common characteristic. 1 4 Perhaps all in one group will be about the same size or about the same color, or have similar shape or the same number of holes. Try to get all the buttons into no more than five such groups. Now sort each of these groups into three or four subgroups in which all the buttons in any one pile are more like one another than they are like the remainder of the group.

Devise a name to describe the buttons in any one of these subgroups. The name should be as brief as possible but should clearly distinguish that group from other groups with different names. The name "small, round, colored, one-hole" could describe buttons in one group while "small, round, colored, two-hole" could identify those in another.

Now draw a classification tree, similar to the one shown, which shows the relation of the buttons in any one of your last groupings to the others.

With the same total collection, follow an entirely different system until you have the buttons sorted again into groups whose members are very much alike. Name these new groups as you did before and construct another tree. $\boxed{16}$ Do you have the same number of groups as you did in the first system? $\boxed{17}$ Are the names of your latest groups the same as those of the first assignment? $\boxed{18}$ Are some of the buttons now in a different group than they were before? Do you feel that your new system is better than the other?

[19] How many different systems can you devise for sorting this collection? [20] Is any one of them similar to the system that would be used in a store?

Buttons, marbles, coins, stamps, or hardware may be used for a classification exercise. Perhaps there is some collection in your home or in the school shop or your science lab that you could sort into more useful classifications. [21] Did you find specimens that were imperfect for some reason, or could not be used for their primary purpose? [22] If so, did you keep them or discard them? There is a story of the tidy little old lady who had all of her possessions so neatly arranged that she even had a box marked "Pieces of string too short to use."

The more refined our system of classification, the greater the likelihood that there will be contradictions and exceptions. There are buttons of such color that they might be placed in any of several color categories. Expect to face reversals in your attempts to classify objects, but do not be discouraged. In almost every filing system there is a folder marked "Miscellaneous." [H, I, J, K]

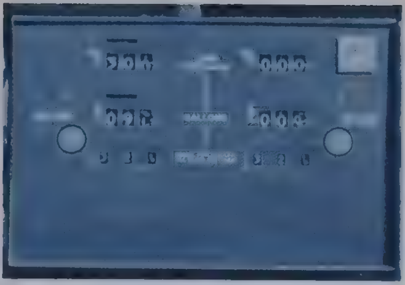
Measurement

We have begun to make observations as we have investigated things around us. In our study of the cube we recorded numbers because it was only when we were able to compare the number of blank faces with the number of faces having a spot that we could reach any conclusions about the marking of the cube.

In our classification practice we made no attempt to be *quantitative*, that is, to take measurements. However, our observations will be more precise, less subject to bias, and less likely to be misinterpreted if we make them quantitative. For example, a button you may have classified as "small" could have been grouped by another person with those called "medium." The words "small," "medium," and "large" are relative terms and depend on the standard with which they are compared. Therefore, standards of measurement are very important in science.

Probably the first measuring rod for length was the human arm. The distance from fingertip to elbow was called a cubit. Smaller distances were measured with the hand. When the fingers were stretched as far apart as possible, the distance from the tip of the little finger to the tip of the thumb was called a span.

The balance is the oldest instrument for measuring weight and mass. People in ancient times used standards made of metal or stone in special shapes such as cubes, flat rectangles, models of ducks, crouching lions, and reclining rams.



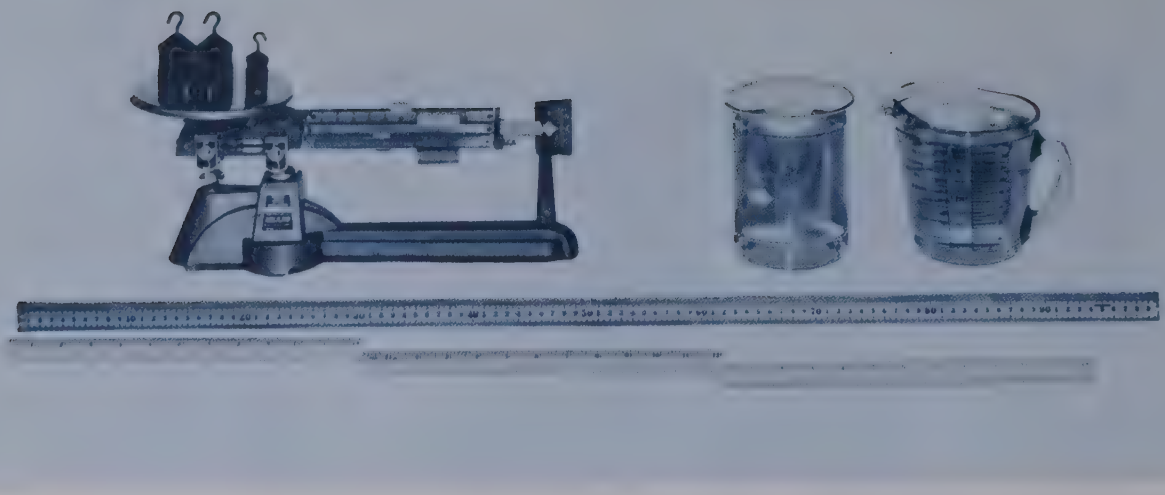
Measuring devices play an important part in our daily activities as well as being scientific tools.

You are familiar with a system of measurement commonly known as the English system. This is used only in English-speaking countries and in a few years the United States may be the only major nation that still uses it. In this system the unit for measuring length is the yard, which is equal to three feet, or thirty-six inches. The unit of volume in the English system is the gallon. (In the countries of the British Empire, however, the standard gallon is not the same quantity as the standard gallon of the U.S.) The gallon is divided into four equal parts called quarts. Quarts, in turn, are divided into two equal parts called pints. The unit of *mass* (weight) in this system is the pound. A pound is divided into sixteen equal parts called ounces.

1–5 THE METRIC SYSTEM

In the late eighteenth century the variety of measurement systems in use and the inconsistent bases for the units were hampering both the work of scientists and the interchange of information between scientists of different nations. Devised by a group of scientists in France, a system was adopted by the International Bureau of Weights and Measures in 1875. This system has three major advantages: (1) it is universally used in scientific work and is the language of scientific measurement throughout the world; (2) it is a decimal system compatible with our base-ten number system; (3) it has consistent relationships between the units of length, volume, and mass. This system is known as the *metric system* (metric means "measurement"). We will use this system in this course.

In the metric system the unit of length is the *meter*. It was chosen for its convenient length and was defined to be one ten-millionth of the



Units of length, volume, and mass exist in all systems of measurement.

distance from the equator to the north pole at sea level. Since the art of measurement was not so highly developed in 1875 as it is now, and since the earth has never actually been measured directly, the meter is only approximately this length. To guarantee that this unit could be faithfully reproduced, two marks were inscribed on a rod made of an alloy of platinum; the distance between these two marks was established as the standard meter. The original rod is kept in a vault in Sèvres, France, and copies of this standard, called secondary standards, are kept in many countries. One such copy is in the Bureau of Standards in Washington, D.C. The latest method of describing the length of the meter is in terms of the wavelength of light from the element krypton. If the original rod were lost or destroyed, it could be reproduced to an accuracy of better than one part in 100,000,000 from this information.

The meter, which is slightly longer than the English yard with which you are familiar, is divided into smaller, sometimes more convenient, units of length. Latin prefixes are used with the word *meter* to indicate the fraction of a meter that each unit represents. A decimeter (dm) is one-tenth of a meter; a centimeter (cm) is one-hundredth of a meter; and a millimeter (mm) is one-thousandth of a meter. A useful item of information is the fact that 2.5 centimeters are about 1 inch.

Other prefixes, derived from Latin and Greek words, are used to designate larger and smaller units. The units of the meter are summarized in Table 1–2 along with the proper abbreviation, the multiplying factor, and the exponential notation.

Other units have been designated to describe very large or very small quantities, such as 10^9 or 10^{-12}, but we will not have occasion to use

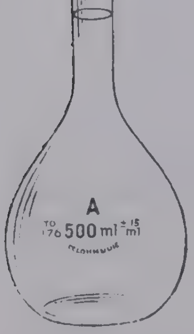


Table 1-2

Prefix	Abbreviation	Multiplying factor	Exponential notation
mega-	(none)	1,000,000	10^6
kilo-	km	1000	10^3
deci-	dm	$\frac{1}{10}$ or 0.10	10^{-1}
centi-	cm	$\frac{1}{100}$ or 0.01	10^{-2}
milli-	mm	$\frac{1}{1000}$ or 0.001	10^{-3}
micro-	μm	$\frac{1}{1,000,000}$ or 0.000001	10^{-6}

The units of the meter.

them in this course. Actually there are very few people who have any real idea of what the actual number of things represented by one million would be.

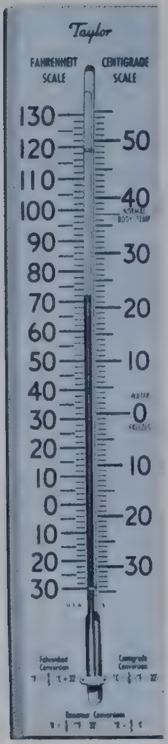
As stated earlier, the metric system has the advantage of consistent relationships between the units of length, volume, and mass. The standard unit for measuring volume in the metric system is the *liter*, which is approximately equal to one quart in the English system. The liter is defined as the volume of a cube one decimeter on each side (1 dm^3, or one cubic decimeter). The same prefixes are used with the liter as with the meter. A deciliter (dl) is one-tenth of a liter; a centiliter (cl) is one-hundredth of a liter; and a milliliter (ml) is one-thousandth of a liter. It is convenient to know that about 20 drops of water from a medicine dropper are required to make one milliliter. Other liquids may have different-sized drops.

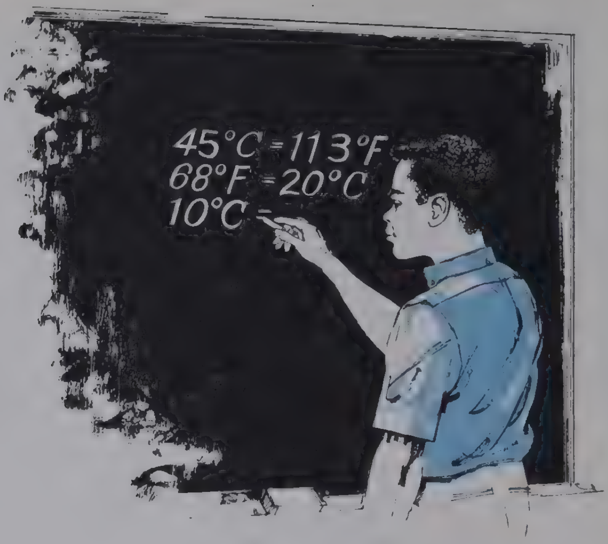
You may have noticed that one cubic centimeter (1 cm^3), which is also equal to one-thousandth of a cubic decimeter, would have the same volume as one milliliter.

The third fundamental unit in this system is the *gram* (g). It is defined as the mass of one cubic centimeter of water. The same prefixes we have already studied may be applied to multiples or fractions of a gram. Thus, a decigram (dg) equals 0.1 gram; a centigram (cg) equals 0.01 gram, and a milligram (mg) equals 0.001 gram. It is convenient to use the kilogram (kg), which equals 1000 grams, for measuring large masses. A kilogram weighs slightly more than a two-pound mass.

The gram is a measure of mass. Although the word *mass* may be strange to you, the concept of mass is not. We will discuss it in Chapter 2 when we begin to work with and measure the masses of objects. Mass is related to weight as you have known it and used it.

To assist you in imagining the magnitude of these new units, let's consider a few comparisons. Most sugar cubes measure approximately one centimeter on a side; therefore, one cubic centimeter (1 cc, or 1 cm^3)





is about the size of a sugar cube. One cubic centimeter, or one milliliter, of water has a mass of one gram. It will be most helpful to you throughout this course if you fix this value in your mind. If this is true, then what is the mass in grams of one liter of water? Remember that one liter is equal to 1000 milliliters.

One more unit of measurement must be mentioned. It is the temperature scale called the *Celsius* (C), or *centigrade*, scale. It is used in scientific work in preference to the *Fahrenheit* (F) scale. The zero point on the Celsius scale is placed at the freezing point of water and 100°C is placed at the boiling point of water. Thus 0°C = 32°F and 100°C = 212°F. More will be said about temperature and its measurement in Chapter 5.

1
5

For practice in using the metric system of measurement, reclassify some buttons into divisions based on their diameters. Those buttons smaller than 8 mm might be classed as "small," those from 8 mm to 14 mm in diameter could be grouped together as "medium," and those 15 mm or larger could be classified as "large." Since each of these groups has such a large range of size, it might be more useful to make more size groups with the members more nearly the same diameter. In order to be used interchangeably as a fastener on a garment, two buttons should be within one or two millimeters of the same size. 23 In a system such as that suggested above, in which group would you place a button with a diameter of 14.5 mm?

If you have a relatively sensitive balance available, you could also sort the buttons into their initial groups by weighing them. Two members of any group might differ by not more than two- or three-tenths of a gram (0.2 to 0.3 g). [24] How closely do members of one of these groups compare in diameter?

If this new system of measurement seems strange to you, take a few measurements using the metric units and try to think of objects in your daily life in terms of meters, liters, and grams. Very soon you will be as skillful in using the metric units as you are now in using English units. If you become discouraged, just remember that you have had long practice with the units of the English system. Similar practice with the metric units can make them easy to use. Furthermore, all scientific measurements are in metric units. [L, M]

1–6 INDUCTION AND DEDUCTION

After scientific information has been collected it must be organized in a useful manner. Then the data must be analyzed in order to obtain true scientific knowledge. At this stage reasoning is most important. The scientist looks for patterns in his data which suggest generalizations about what he has found. This form of reasoning is known as *induction*. It requires much data, and experiments must be repeated many times. You used induction to decide the probable pattern on your sugar cube. Inductive reasoning is important in formulating general laws in science.

Many centuries ago astronomers may have observed that the moon passes through the complete cycle of its phases approximately every 29 or 30 days. Thus it appeared to revolve around the earth in a regular way. From this general pattern of motion, careful observers were able to predict specific information about the moon's appearance at any desired time in the future. It was even possible to predict the occurrence of eclipses. This predictive process is *deduction*. It is the result of thinking more than of experimenting. When we attempt to predict a particular event or observation from our knowledge of a general theory, we are using deductive reasoning.

Both induction and deduction are useful tools in science. You have used both many times, probably without being conscious of what you were doing. In our very first experiment we observed the behavior of two samples of bread. All conditions, so far as we know, had been kept the same except that one sample was moistened while the other was not. From this we may have been able to detect a difference in the growth of molds. If the bulk of the data gathered by your group showed that a greater mold rate occurred with moistened bread, you might inductively arrive at the conclusion that the rate of mold growth is influenced by the

What is science? It has many aspects.

We must observe and record . . .

presence of moisture. This one conclusion could then be used to deduce information about what you would expect to happen in other circumstances.

You will call on your powers of reasoning every time a problem is encountered in the study you are beginning here. Hopefully, you will carry this same process into your attack on other problems of life. Notice the manner in which we have approached the solution of our first two problems. We have gathered evidence under carefully controlled conditions until we felt we had enough good information to make a reasonable conclusion possible. Then, by applying our minds to the collection of data, we arrived at the best answer we could get. If our experiment had permitted, we would have tried to predict other consequences of the conclusion. [25] If it had happened that in the bread-mold experiment the unmoistened sample was always on the left and the moistened sample was always on the right, could you honestly say that it was the moisture that had caused any differences that may have been noted? Common sense would probably tell us that this position difference could not be responsible for any such change, but it is important to realize that common sense is the accumulation of all the conclusions we have drawn from our observations and other learnings throughout life. It is possible to live all of life with a set of common-sense notions that would not hold up under careful scientific testing.

Think about the statements and "facts" you encounter before you accept them. Take a critical approach; keep in mind that some sources of

measure and calculate . . . organize and examine . . . read and interpret . . . for there is much to learn.

information are more reliable than others. Much of what we know is based on conclusions drawn by the finest minds and on centuries of observations made by many persons. To a large extent, this offers us the authority we need to accept almost everything we are expected to learn. But we should never blindly accept any "fact." Allow yourself the pleasure of evaluating "facts" after you have gathered enough data so you have something to work with. Develop your powers of observation; learn to organize and keep your information in a useful manner; think, both from the specific to the general (induction) and from the main ideas back to the individual details (deduction). These are the mental methods of the scientist.

You might think of science as a game in which you are competing with the unknown. You win if you have increased your knowledge and your understanding. You must observe and record, measure and calculate, organize and examine, wonder and guess. Then you must check yourself, the things you are observing, information written or published on the subject; anything, in short, that will help you to understand what you have done and point the way for further investigations. Science is a subject that has captured the attention of some of our finest minds. Its rewards have given us a way of life beyond the imagination of those early scientists who first began to wonder about the sun, the stars, the earth, and life itself. And there is so much to learn, far beyond what you will encounter in this course. There is an entire universe challenging you to study, to discover, and to understand its order and beauty.

TAKING INVENTORY

Science is a human activity that enables us to find meaningful answers to many of our questions. Both direct and indirect observations are important in our pursuit of answers. Accurate records of each experiment furnish evidence from which we can draw conclusions. Wise use of controls is necessary if results are to be of value.

Tools such as grouping (classification), measurement, and reasoning require practice if we are to gain skill in their use. As we sharpen our senses, become increasingly aware of our surroundings, and develop a greater ability to ask questions, our daily activities will become more exciting.

FURTHER EXPLORATIONS

A. On page 3 six different questions were asked about various factors that might influence the growth of mold on bread. What hypothesis would you offer as a predicted answer to each question? What sort of experiment can you design to test each of your hypotheses? What additional questions can you raise about factors that might influence growth of the mold? Could you design an experiment with appropriate controls that would test more than one factor at a time? Select one of these questions or devise a question of your own regarding the molding of bread. Advance a hypothesis and plan an experiment to test this hypothesis.

B. After performing the bread-mold experiment, can you criticize it and make it into a tighter, better-controlled experiment that will produce better data?

C. The nature of our experiment with bread mold produces only a yes-or-no answer to the question. With the addition of other elements, however, we might reach a more refined answer. Set up a more elaborate experiment using several jars and moisten each piece of bread with a different amount of water; or with several jars and several slices of bread, try to learn how the rate of molding is related to temperature by placing each jar in an environment with a different, fairly constant, temperature.

D. If the cube used to obtain the information given in Table 1–1 had actually had two of its six faces marked, the marks could have been arranged in only two possible patterns. The two marked faces might have been side by side or they might have been on opposite sides. Does your experimental procedure give any useful information about this question? If it does not, can you devise a method of indirectly observing in order to answer such a question?

E. Assume that a cube you examined showed evidence of having three spots.

a. In how many *different* arrangements might you place a mark on each of three faces of a cube? Could you determine, with your original apparatus, which method had been used? Can you determine this with certainty by *any* indirect method?

b. In how many *different* arrangements might you place a spot on four faces of a cube?

F. Assume that you are a marketing-research director for a sugar-refining company. You have determined the best size to make cubes of sugar so that they correspond to the measures people are accustomed to using with granulated sugar. Now you wish to find the most efficient and economical method of packaging them for marketing.

If 180 cubes are required for a one-pound package, into how many rows, how long and how high, would you assemble these cubes so that the package would have the smallest outside surface? How many different kinds of packages could be devised if you did not have to restrict the size of the outside surface? Devise a system for studying the problem. How many of the possible arrangements must you check by calculation before you may make a definite statement about the most compact arrangement?

G. In a series of 100 observations of eight cubes, the following information was obtained.

A spot was observed on each of the eight cubes 4 times.
A spot was observed on seven of the eight cubes 20 times.
A spot was observed on six of the eight cubes 35 times.
A spot was observed on five of the eight cubes 16 times.
A spot was observed on four of the eight cubes 20 times.
A spot was observed on three of the eight cubes 5 times.

What can you conclude about the markings on the cubes? Can you draw conclusions about the markings on any one cube?

H. In art class you may have learned that we classify colors. One classification is termed the primary colors. It contains three classes: reds, blues, and yellows. In science class you probably learned that there are seven colors in the spectrum (rainbow). At a paint store you may find a machine with which you can mix hundreds of colors. Why, then, do we say there are only three primary colors? Are there really only seven colors in the spectrum? Examine a spectrum carefully and see if you can distinguish where one color stops and another color begins. Can you suggest a better classification for the colors of the spectrum? Are these the same colors you observe when sunlight is reflected from a thin film of oil on water?

I. For further practice in devising systems of classification, select one or more of the groups suggested here and see how many ways you can sort them: leaves, sea shells, rocks and minerals, clouds (a collection of photographs of cloud formations may be used).

J. We classify our symbols such as letters, words, and numerals. How many systems can you identify for classifying words? for classifying numerals and numbers?

K. If you have made a hobby of collecting stamps, coins, or other articles, examine your collection to determine the basis for the classification system you use. For your collection, construct a "classification tree" similar to the one shown for buttons. Is there another, more useful way you might group your individual items?

L. Is there any relationship between the classification of objects and the measurement of length, volume, and weight? Explain your answer.

M. Using a medicine dropper, measure the number of drops of tap water required to make one milliliter. You may need to drop out ten or twenty times this much in order to obtain an accurate measure. Use the same medicine dropper to determine if the addition of such material as soap or a detergent will change the average size of the drop. Would your dropper produce the same size drop of gasoline, milk, engine oil, or rubbing alcohol? If the dropper had a blunt end or had a very fine point, would this affect the drop size? Would a plastic dropper affect the size of drop? Choose some question about this system of measurement, make your guess of the answer, and then devise a method of attack (proceeding with the experiment) to assure useful answers. You may need to design and make some of your own equipment to guarantee that you draw reliable conclusions.

FURTHER READING

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———, *Probability and Chance.* Webster Division, McGraw-Hill Book Company, New York, 1963.
An excellent booklet with many exercises to introduce the student to a challenging area of science.

———, and WILLIAM H. GLENN, *The World of Measurement.* Webster Division, McGraw-Hill Book Company, New York, 1961.
Provides a good discussion on the history and growth of the science of measurement with numerous exercises to give the student practice in many areas of this important scientific skill.

———, and WILLIAM H. GLENN, *The World of Statistics.* Webster Division, McGraw-Hill Book Company, 1961.
Provides an easy introduction to a difficult field important to the science student. Many excellent illustrations and exciting exercises to give experience in the use of statistical data.

Patton, A. R., *Science for the Non-Scientist.* Burgess Publishing Company, Minneapolis, 1962.

Gives a good background in what science is. Should be read by every student of science.

Ruchlis, Hy, *Discovering Scientific Method.* Harper and Row, Publishers, New York, 1963.

Pages 1–31 give an excellent introduction to the methods of science. Well written, with many examples of how scientists go about solving problems.

Youden, W. J., *Experimentation and Measurement.* Scholastic Book Services, New York, 1963.

A paperback which students will find most helpful in understanding measurement as a tool of science.

2

Investigating the Structure of Matter

2–1 Early Science
2–2 Hypotheses to Explain the Structure of Matter
2–3 Checking Our Hypotheses
2–4 Assembling Evidence About the Structure of Matter
2–5 Packing of Particles—Another Clue to Structure

2–1 EARLY SCIENCE

For centuries man has proposed many theories to explain the structure of matter. Let's examine some of these theories and see how man has used them to attempt to answer these questions: What might be the nature of matter? What general characteristics can we find that are common to the matter of which all things are made?

Our present theories owe much to the rich heritage of scientific thinking that began in Greece centuries before the birth of Christ. Although the names and works of many men have been lost during the passage of time, there are still a number of philosophers who are remembered today for their contributions to the theories of matter.

One of the first of these early scientists who developed a significant theory of matter was Thales, who lived about 600 B.C. in the seaport of Miletus, on the western coast of present-day Turkey. He believed that water was the fundamental material of the universe. Heraclitus of Ephesus, a seaport near Miletus, said in 500 B.C. that fire was the basic component of all matter. Heraclitus also believed that all matter was in a continual state of change and nothing was really permanent.

Empedocles, born in 495 B.C. in a Greek community in Sicily, combined water, fire, air, and earth into his theory of matter. He claimed that under the influence of Love (harmony), these elements united into an organized whole. At other times, when Strife (discord) became strong, the elements were torn apart. Since earth, water, and air were examples

600 B.C.
Thales believed that water was the fundamental material of the universe.

500 B.C.
Heraclitus thought that fire was the basic component of matter.

495 B.C.
Empedocles thought there were four basic elements: fire, water, air, and earth.

460 B.C.
Democritus advanced a theory of tiny particles.

200 B.C.
Archimedes discovered the relation of mass to volume for matter.

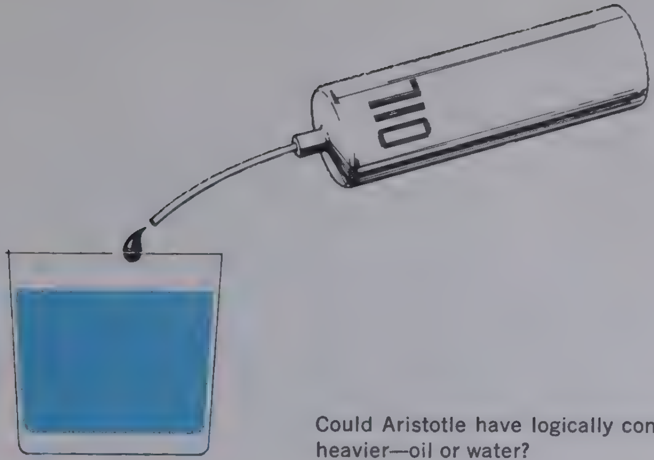
350 B.C.
Aristotle refined a view of earth, air, water, and fire as the four elements.

TODAY
The Fabric of the Universe:
Matter and Energy

of each of the three states of matter (solid, liquid, and gas), and since fire is a form of energy, part of Empedocles' theory is in accord with our modern views.

Democritus (460–370 B.C.) was born in Abdera, in northeast Greece, but he traveled and studied in Egypt and elsewhere. He advanced the theory of his master Leucippus that matter is made of tiny particles. Democritus called these particles *atoms*, from the Greek word *atoma* which means "not divisible."

Aristotle (384–322 B.C.), the great intellectual giant of Greek culture, favored and further refined the idea of Empedocles that matter consisted



Could Aristotle have logically concluded which was heavier—oil or water?

of four elements (earth, water, air, and fire). Because of his great reputation as a thinker, Aristotle's teachings about the nature of matter were accepted by both philosophers and alchemists. During the Middle Ages, his theories became entwined with Church doctrine. To question or attack these beliefs was to attack the Church itself, and many thinkers who dared to doubt Aristotelian theories in public were punished or put to death.

Aristotle's influence on scientific thought and development was felt for some 2000 years. He made many remarkably acute observations, and he drew conclusions from them that were to result in much scientific progress. Today, our criticism of his work is not for his contributions to science and knowledge, but for those writings and teachings that were based more on how he thought things ought to be than on careful observation of how they actually were. The following anecdote is an example. As part of his teachings, Aristotle established a scale of perfection for all things. He put living things at the top of the scale, and classified the male human being as the most perfect living thing. Women were classified as less perfect than men. One reason for this, according to Aristotle, was that women had fewer teeth than men. Much later, someone finally decided to investigate the idea and found that women have the same number of teeth as men. Unfortunately, the conclusion at the time was not that Aristotle had been wrong, but that the nature of women had changed! It is possible that Aristotle's idea about fewer teeth in women came from the fact that the female horse normally has four fewer teeth than the male. But we have come today to realize that, if we want reliable information on any subject, we should try to get it from direct observation.

Like Aristotle and his theory of the scale of perfection, the other early scientists and philosophers did not test their theories by doing the kinds of observation and experimentation we have been doing. They felt

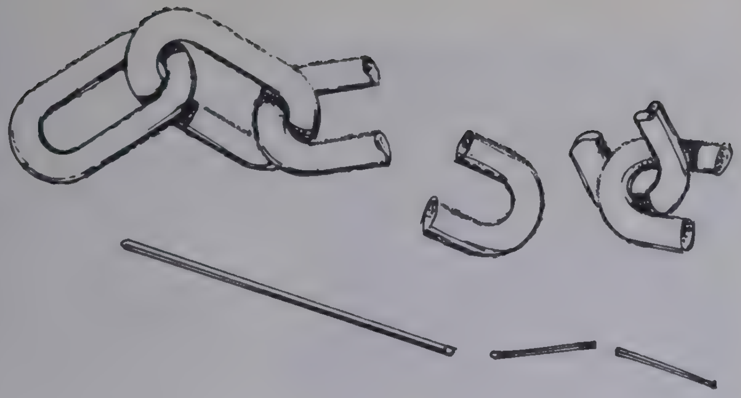
Can we fit celestial bodies into a theory of matter?

that it was sufficient to think about the problem and decide what the logical answer should be. And since these early philosophers were so widely respected, others did not doubt them and did not think it necessary to check their statements against nature. Of course, evidence about the underlying structure of matter is hard to determine since it cannot be directly observed and must be studied by indirect methods. Thus, it is not hard to understand why the early thinkers preferred to do their experimenting in their heads.

Any theory, however, that stimulates investigation and results in fruitful searching has merit, and, incorrect as they were, these early theories of the structure of matter prompted thinkers to ask questions. As Francis Bacon, an English scientist who pioneered the scientific method in the sixteenth century said, "Truth arises more readily from error than from confusion." Thus even these early and currently unacceptable theories gave direction to the pursuit of knowledge.

2–2 HYPOTHESES TO EXPLAIN THE STRUCTURE OF MATTER

Matter is all around us. It includes water, sand and stone, plants, and animals. [1] Does it also include the sun, other stars, the planets, and other objects that we know are not part of the earth? [2] Does it include waves on water and the wind? The dictionary defines matter as "whatever occupies space." Just what does this mean? Let's attempt to construct a mental model for the structure of matter to explain what we observe and allow us to make predictions about further observations.

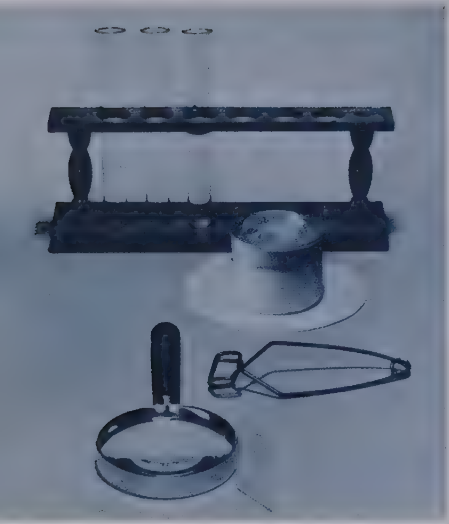


Is matter clearly particulate or clearly continuous in appearance?

Is the railroad track continuous or particulate? Is the fence continuous or particulate?

From the many descriptions of matter, we will try to limit our investigations to three possibilities. Each of these is a guess, or hypothesis, which should be checked by experimentation. First, we might assume that *matter is composed of tiny particles.* Second, we might guess that *matter is continuous*—with no distinctive particles. We might use a chain and a solid wire to illustrate these ideas. Both the chain and the wire make up continuing objects. Regardless of how short we cut the wire, the pieces are still wire and differ from one another only in length. But if we separate the links of the chain, it is no longer a chain, for it does not have the properties of a chain. The wire, in this example, appears continuous, whereas the chain is made of particles called links. Another comparison may be made with a dotted line and an unbroken line. The dotted line may be considered to be *particulate* (composed of particles) while the unbroken line is continuous. But since both the chain and the dotted line have certain characteristics in common with the wire and the unbroken line, we come to the third possible hypothesis: that *matter exhibits both particulate and continuous properties.*

Often the way things appear depends on how we look at them. From high above, a stream of cars on a highway looks like an unbroken band. At closer range we see that it is made up of individual cars. However, at the even closer range of the pedestrian, we know that we cannot safely break through this continuing strip until the traffic signal changes and the cars stop. Whether traffic appears continuous, particulate, or both may depend on how we look at it—upon our point of view. Likewise, whether matter appears continuous, particulate, or both may also depend on our point of view.



Move the test tube back and forth in the flame rather than applying heat to one place. Never point the open end of the test tube toward yourself or anyone else while heating it.

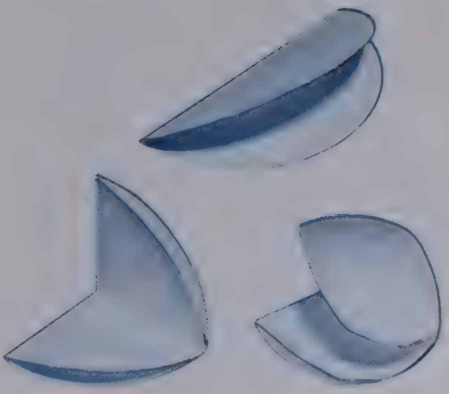
2–3 CHECKING OUR HYPOTHESES

The Forms of Water

Let's begin our study of matter by observing water in its different forms. We most often use water in liquid form, but we also use ice, the solid form of water. The gaseous form of water, steam, occurs whenever we boil water. By giving careful attention to the three forms of water, we may be able to gain insight into the structure of matter.

[2] [1] For this experiment you will need a test tube, water, a magnifier, a glass slide, a test tube holder, a source of heat, a small metal can, table salt, and enough crushed ice to fill the small metal can. Be sure to keep dated records of all observations in your laboratory record book. Without magnification, water appears to be continuous. Can it be magnified sufficiently to show evidence of particles? Put a drop of water on the glass slide and examine it with a magnifier. [3] Do you find evidence of particles or does the drop of water still seem to be continuous? We must not overlook the possibility that the power of our magnifier may determine our ability to see particles.

Heat the water in the test tube until it boils. Fill the metal can with crushed ice, place it in the steam rising from the boiling water, and observe the tiny droplets which condense on its surface. [4] Are the droplets the same size? [5] Is it possible that some of them are too small to be seen? You may have noticed that when water is left in an open container for a period of time it disappears, or *evaporates*. [6] Does this



The steps in folding a piece of filter paper.

Heating the sulfur sample.

Pouring the molten sulfur into the filter paper.

change help us decide whether water is continuous? [7] How would it behave if it is continuous? [8] How would it behave if it is composed of particles?

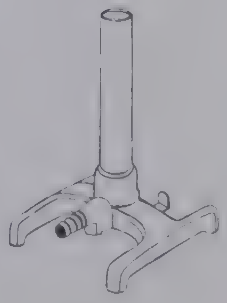
Sprinkle the ice in the can liberally with salt. Breathe gently on the outside of the can until frost forms. Examine the frost with the magnifier for evidence of particles. You may want to look at the ice with the magnifier and compare its appearance with the frost. Both frost and ice are solid forms of water. [9] In what ways do they differ?

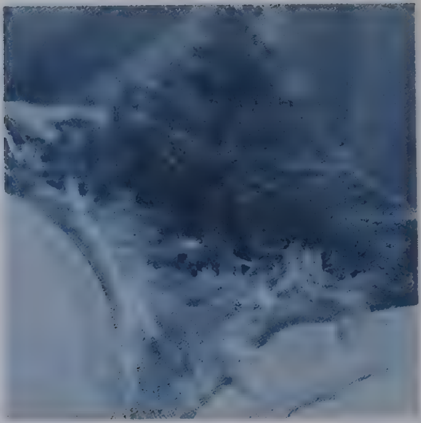
The Forms of Sulfur

Sulfur is another common substance. It exists as a solid at room temperature, but it can be changed into a liquid and eventually into a gas by heating. It is an interesting material because it assumes different forms even in the solid state. We can gather many clues to the structure of matter by a study of this *element*. (An element is a substance that cannot be separated into any other substances.)

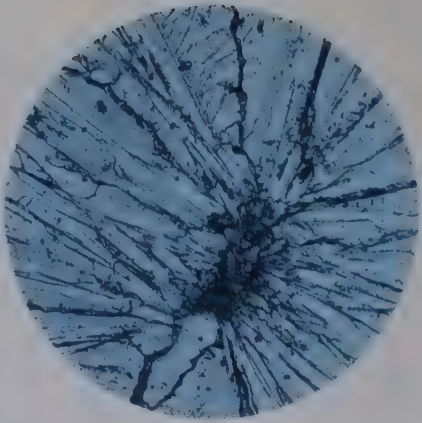
For these experiments you will need a piece of roll sulfur, a small test tube, a test tube holder, a source of heat, a glass slide, filter paper, a funnel, a stirring rod, a magnifier, a beaker of cold water, and carbon disulfide. Be sure to keep dated records of your observations. Try to relate each individual experiment to the main idea under investigation. In these activities we are still checking our models of the structure of matter.

Examine the broken faces of a piece of roll sulfur and note the general





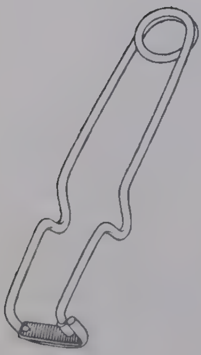
Have you succeeded in preparing a crystal sample that looks like this?

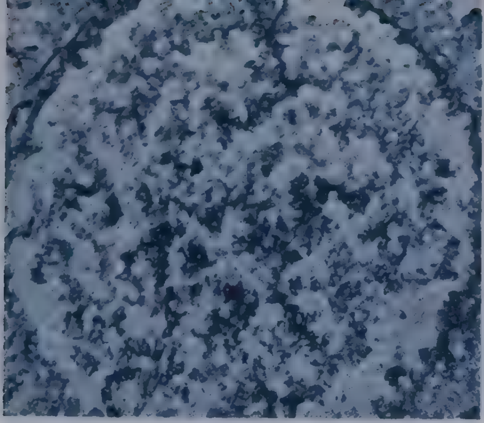


Microphoto of prismatic sulfur.

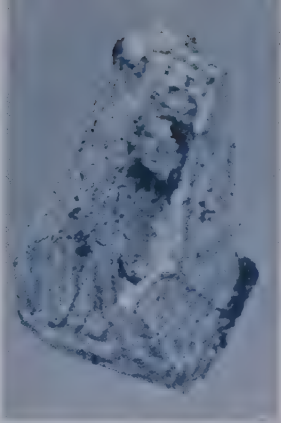
characteristics of the element. When you are satisfied that you are familiar with the solid sulfur, place a few chips of it in a small test tube and heat the tube *gently*. It is important to keep the heat low enough so that the sulfur does not lose its yellow color. As soon as the sulfur has melted, pour the yellow liquid onto a piece of filter paper which has been properly folded and placed in a funnel. When the surface of the molten sulfur has cooled sufficiently to form a crust, remove the filter paper from the funnel and spread it open to expose the region where both solid and liquid sulfur come together. Examine the resulting sulfur with a magnifier. If you have good luck, you will be able to see long, slender needles. We call these regular patterns *crystals*. This crystal form of sulfur is known as *monoclinic*, or *prismatic*, sulfur. Sketch an example of this monoclinic sulfur in your record book for comparison with other forms you will be preparing. $\boxed{10}$ Can you explain why the liquid sulfur forms these crystals upon cooling? Perhaps we now have evidence that there are very tiny pieces of sulfur which fit together in an orderly way during the cooling process to produce this characteristic pattern. Label and date this sample of sulfur, and save it for examination during a further study of crystals in Chapter 4.

Sulfur can be taken out of its solid state by another process known as *dissolving*. Although sulfur does not dissolve in water, it dissolves rather readily in the compound carbon disulfide. (**CAUTION:** Carbon disulfide is dangerously flammable and its vapors are toxic. Be sure to perform these experiments in a well-ventilated room where there are *no* open flames!) Crush a pea-sized piece of roll sulfur and put it in a test tube containing 5 ml of carbon disulfide. Stir the mixture with the stirring rod until the sulfur completely dissolves. Evidence of the sulfur

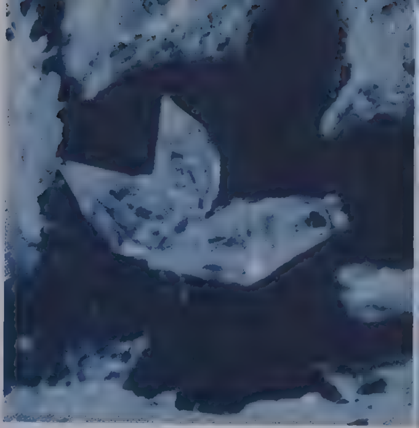




Rhombic sulfur crystals on a watch glass.



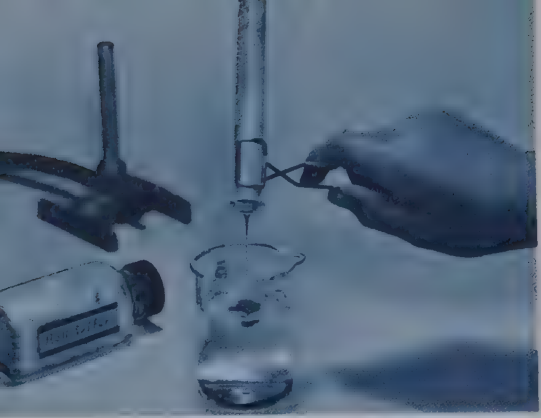
Rhombic sulfur crystals under a magnifying glass.



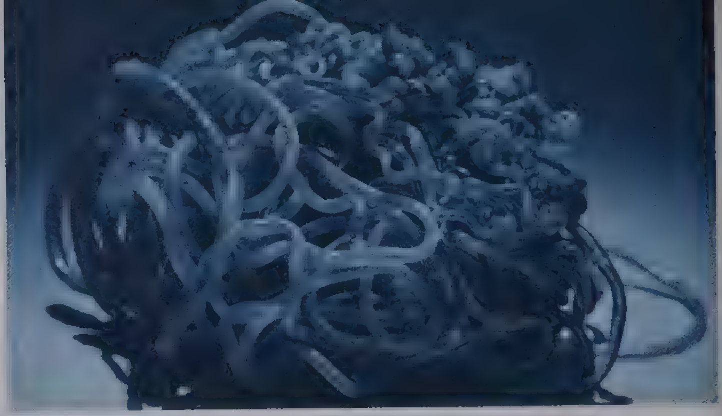
Microphoto of rhombic sulfur crystals.

can be seen in the pale yellow color of the solution. Transfer a drop of this solution to a clean glass slide and watch with a magnifier while the carbon disulfide evaporates. Compare the shape of the sulfur crystals that form with those formed by slow cooling in the filter paper. Some of these crystals should be diamond shaped; they are known as *rhombic* crystals. Allow the remaining sulfur-carbon disulfide solution to evaporate slowly for several days. Large rhombic crystals of sulfur will form. You will need one such large crystal for a later experiment on density.

There is still another form of sulfur we should investigate. Half fill a test tube with crushed sulfur. Heat the sulfur until it turns dark. Sulfur vapors may catch fire at the mouth of the test tube and burn with a pale blue flame. (**CAUTION:** The gas formed by burning sulfur is irritating and unpleasant. Do not inhale it, since it is harmful to nasal tissues and lungs.) During the heating process, it is interesting to tilt the test tube and note the changes in the ease with which the liquid sulfur moves. Pour the dark, molten sulfur *slowly* into a beaker of cold water. If you heated the sulfur sufficiently, the resulting mass will be in the form of dark-colored, rubbery strings. You will probably find no evidence of crystals. (**CAUTION:** Do not handle the rubbery mass until it has cooled enough to avoid burning your hands.) Although every solid object has a form, the word *amorphous*, which means "without form," is used to describe sulfur prepared in this manner because it does not have a form common and consistent for all such samples. Each of the other two forms had a characteristic structure. When the amorphous sulfur is cool, examine it with the magnifier. Keep this sample, labeled and dated, for later examination. Interesting changes will occur which can add to your growing knowledge of the structure of matter. [A]



Preparation of amorphous sulfur.



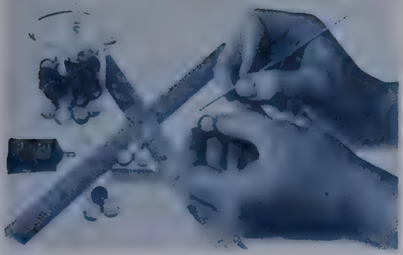
A close-up of amorphous sulfur.

2–4 ASSEMBLING EVIDENCE ABOUT THE STRUCTURE OF MATTER

We have found that water exists in liquid, solid, and gaseous forms and that these forms seem to be dependent upon temperature. Water appears to be a continuous substance, but the fact that it produces a regular pattern (crystals) under certain conditions and disappears in the air as it boils leads us to believe that it may also have particulate properties.

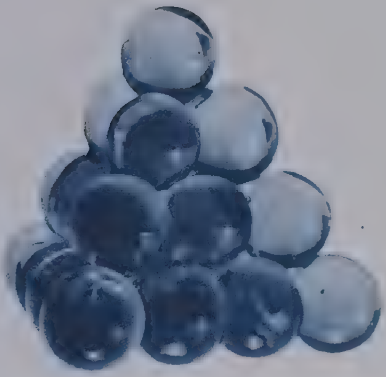
Sulfur exists as a solid at room temperature but can be changed into a liquid and finally into a gas when heated. In the solid state, it forms either prismatic or rhombic crystals, thus showing evidence of regular patterns. The semisolid, rubbery form shows no crystalline properties that are evident to the unaided eye. 11 Could a microscope help you to see evidence of crystalline structure?

12 Which of our proposed theories for the structure of matter can best explain the observations we have made about water and sulfur? 13 Both substances appear to be continuous, but if they really are continuous, would the changes in form be possible? If they are composed of tiny particles, then it is easy to assume that these particles can be arranged in different ways to produce the forms we have seen. Crystallization into apparent patterns took place, you will recall, when the cooling process occurred quite slowly. The amorphous form was produced by a rapid cooling process. If our theory of particles is correct, we may say that, during this rapid process, the particles bunched together so hastily that no regular pattern could form. We can compare the rubbery strands of amorphous sulfur to a handful of straws or spaghetti that we quickly tossed into a box rather than carefully stacking them into a neat



Building a tetrahedron.

Constructing a pyramid.



arrangement. If we were to repeat this tossing, the pattern would be different each time. Certainly at this point we must conclude that matter, as we have observed it in water and sulfur, has both continuous and particulate properties. [B]

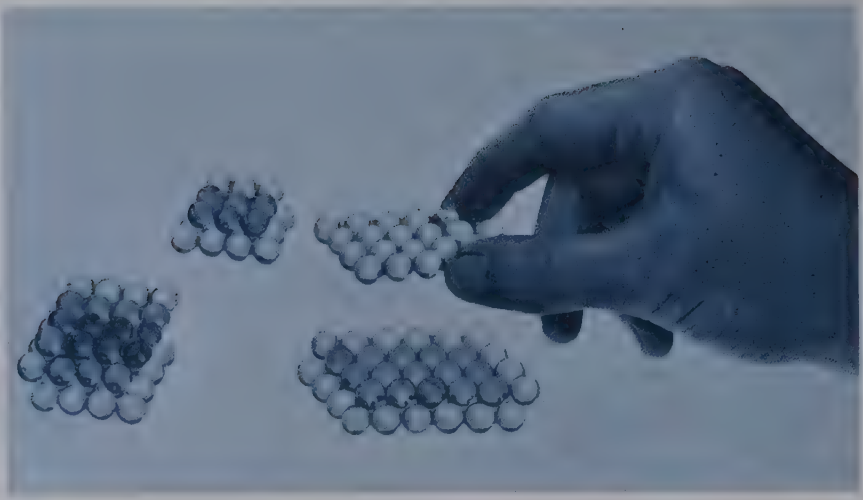
Building Replicas of Crystals

Let's assume that each of the crystals we have studied owes its shape to a unique arrangement of tiny, identical particles. If we can use identical particles which are large enough to see to build crystallike structures, we will have evidence that a particle model is reasonable.

For this experiment you will need 50 to 75 marbles of the same size, quick-drying cement, and a candle. The first model we will build is a very compact arrangement of four marbles. Arrange three marbles in a triangle. Glue them together using a drop of cement between adjoining marbles. Glue one additional marble to the base formed by the three marbles by cradling it in the depression in the top. This model now has the same appearance when viewed from any of its sides. Such a figure is called a *regular tetrahedron*. The word *regular*, as used here, means that each of the faces is like every other face. *Tetra-* is a Greek prefix meaning "four."

Build another tetrahedron like the first one, but use drops of candle wax to hold the marbles together instead of cement. Keep this model for an experiment in Chapter 3.

You may want to build another similar figure with cement having three, four, or five marbles on each side of the base. For purposes of comparison, however, you should build a model having a square base. Use sixteen marbles to make a square base. Build on this base by adding



a marble to each of the small depressions formed where four marbles meet. When this model is thoroughly dry, examine it. [14] How many marbles does it contain? [15] Are each of the triangles that form the sides the same shape? [16] Are the marbles in this model as closely packed as those in the first models? If you build another model in the same manner, starting with a square base with three marbles on each side, you can place the two bases together to form a double pyramid. This figure has eight similar faces and is known as an *octahedron*. *Octa-* is a Greek prefix meaning "eight."

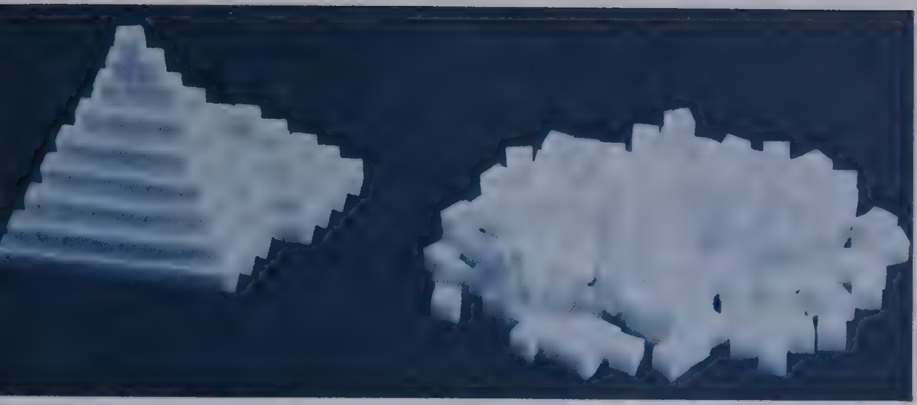
Cement several sets of six marbles into straight rows. After the cement has hardened, assemble these rows into sheets two, three, or four marbles wide. Then use these sheets to form as many different kinds of stacks as you can. Note the characteristics of each different arrangement.

[17] Do any of these marble models resemble either the water or sulfur crystals that you have observed? The fact that identical particles can be arranged to form regular solid figures lends some support to the hypothesis that matter is made of particles. Crystalline matter shows some of the properties we find in the objects we made of visible particles. [C]

2–5 PACKING OF PARTICLES—ANOTHER CLUE TO STRUCTURE

Density

If a crystalline structure represents an orderly arrangement of small, identical particles of a substance, and *if* formlessness represents a lack of order, *then* we should be able to notice some difference in properties



Each pile contains an equal number of cubes. Is the space filled by the two piles equal?

due to the difference in the packing of particles. For example, if a given number of sugar cubes is arranged in the most orderly manner, then they should occupy less space than they would if they were bunched together without order. We can observe this if we empty a one-pound box of sugar cubes and then try to put the cubes back in the box by pouring them in, instead of stacking and packing them. In a stack of carelessly piled sugar cubes, open spaces between the cubes exist. The weight is the same but the volume—the space they occupy—is different.

In order to consider packing of particles quantitatively, we will use a new unit called *density* to describe it. For the present, we will define density as the weight of our sample divided by its volume. In metric units, density is measured in grams per cubic centimeter (g/cm^3), or grams per milliliter (g/ml).

Earlier, we suggested that the rapid cooling of molten sulfur produced the amorphous form because the time was too short to produce the orderly crystalline arrangement of particles. If we can determine a difference in density, it will give support to a particle theory. If we can determine the volume and weight of each of the amorphous and the crystalline types, we may be able to make the comparison. A crystal of sulfur, if it were large enough, could be measured directly, and its volume could be calculated. It is easier, however, to determine the volume of most of the samples of sulfur we can obtain by immersing them in water to determine the space they occupy. If the sample of sulfur is dropped into water in a container, the increase in volume can be measured by observing the rise in water level. We will assume that this increase is the volume of the sulfur sample.

Determining the volumes of sulfur samples.

For this experiment you will need a piece of roll sulfur or a very large crystal, and a mass of amorphous sulfur, both samples large enough to be weighed conveniently. You will also need a 100-ml graduated cylinder for measuring volume and a balance for weighing the sulfur. Weigh each of the two kinds of sulfur as precisely as possible and record these weights in your laboratory record book. Put 50 ml of water in the graduated cylinder, measuring the amount carefully. Immerse the roll sulfur in the water and carefully read the new volume. Subtract the first reading of volume from the second. This difference, which you should enter in your laboratory record book, is the volume of your sulfur sample. Now divide the weight of the sulfur sample by its volume to get the density of the crystalline sulfur—the number of grams in one milliliter—and enter this in your laboratory record book.

Determine the volume of amorphous sulfur by immersing your specimen in water as you did the crystalline sulfur. Calculate the density of amorphous sulfur by dividing its weight by its volume. [18] What does this density indicate about the packing of the particles that may make up each of these sulfur varieties? [19] Is the density of a fresh sample of amorphous sulfur the same as that of a specimen prepared several days earlier?

This section opened with a statement about particle arrangements and their effect on properties of substances. The statement began with an "if" clause and ended with a "then" clause, a common way of stating a scientific hypothesis. [20] Does your evidence about the density of the two forms of sulfur support the hypothesis? [D, E]

Mass and Weight

When we introduced the metric system in Chapter 1, we described the gram and the kilogram as units of mass. Frequently, especially in everyday language, grams and kilograms are spoken of as units of weight. Now that we have mentioned weight in relation to density, we need to clarify the meanings of the terms *mass* and *weight*. Such clarification will aid our continuing effort to understand the structure of matter.

Weight is a force. Specifically, it is a measure of the gravitational force on an object at any given location. Scientists believe that gravity on the moon is only about one-sixth as strong as it is on earth. Thus an object on the moon would weigh one-sixth of its weight on earth. The weight of an object varies slightly even at different locations on the surface of the earth because the gravitational field varies. The mass of an object—the amount of matter in it—does not change unless some matter is removed or added.

Technically, then, we are actually concerned in our study of density with mass rather than with weight. The fact that an object changes weight as it moves through space, from the earth to the moon, for example, does not change its density. That is, the move does not change the relation of mass to volume.

If the density changes, it does so because of a change in the arrangement of the matter within the sample, which has an effect on the volume. You were trying to detect just such a change in arrangement and to use it as evidence for particles when you determined the density of two forms of sulfur. We will now refine our definition of density by saying that it is the ratio of *mass* to volume.

The easiest way to find the mass of an object is to weigh it, since the weight of an object is directly related to its mass. It is quite common to say *weight* when we mean *mass*, and as long as we are making weight measurements under similar circumstances, it is not wrong to describe the masses of objects by using their weights. But since weight in its strictest definition is a force, we should not use the words *gram* and *kilogram*, which describe mass, to describe weight. Instead, we will refer to the earth's gravitational pull on a one-gram mass as one *gram-force* (gf). One *kilogram-force* (kgf) is defined as the earth's gravitational pull on a one-kilogram mass. If you weigh an object on a two-pan balance, you balance the pull of gravity on the object by putting masses in the other pan that provide an equal and opposite pull. When the device is balanced, you know that the mass of the object is equal to the mass it balanced. In other words, by weighing an object using the gravitational field of the earth you can determine the mass.

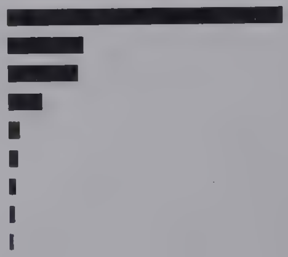
Specific Gravity

Because water is a very common substance, it is often convenient to compare the density of other substances with the density of water. As we learned in Chapter 1, one milliliter of water weighs one gram. Therefore, by applying our definition of density (mass divided by volume), we find the density of water to be one gram per milliliter (1 g/ml). If we divide the density of another substance by the density of water, we obtain a ratio called *specific gravity.*

$$\text{Specific gravity} = \frac{\text{density of a substance}}{\text{density of water}} = \frac{\text{g/ml}}{\text{g/ml}} .$$

Specific gravity is thus a comparison which tells us how many times more dense or less dense a given substance is than water. Since the quantity is a ratio, we do not assign dimension units to it as we do to density. It is called a *dimensionless number.*

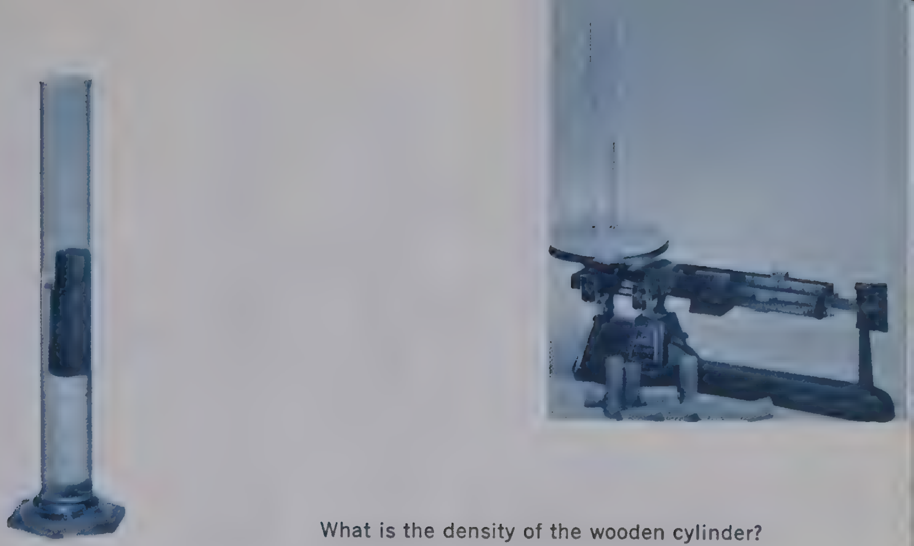
Table 2-1

Substance	Specific gravity	Volume in ml of 100 g	Relative volume of 100 g
Cork	.25	400.0	
Olive oil	.92	108.7	
Water	1.00	100.0	
Table salt	2.16	46.3	
Iron	7.86	12.7	
Silver	10.50	9.5	
Lead	11.35	8.8	
Mercury	13.55	7.4	
Gold	19.33	5.2	

Specific gravity values for some common substances.

Specific gravity allows us to compare quickly the density of two substances. For example, the specific gravity of aluminum is 2.70 and that of magnesium is 1.74. Since both quantities relate to water, we know from examining these figures that aluminum is denser than magnesium and that both are denser than water.

We have already seen that an object immersed in water will cause the apparent volume of the water to increase by an amount equal to the volume of the object. Archimedes, a citizen of the ancient Greek colony of Syracuse in Sicily, determined that there is also a definite relationship between the weight of the object and the weight of the water displaced. In contrast to most thinkers of his time, Archimedes believed in testing



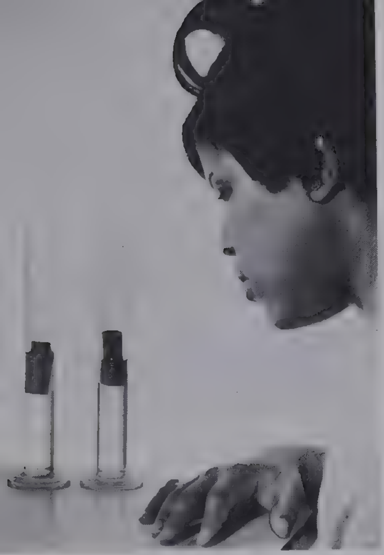
What is the density of the wooden cylinder?

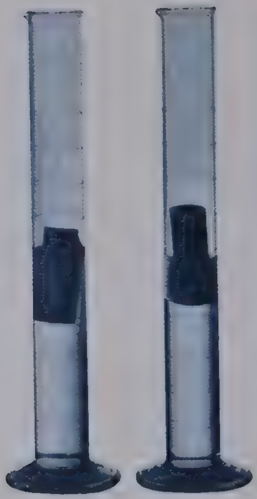
his ideas by conducting experiments. One of his most important contributions was his study of floating objects. Because he was the first to make scientific observations on this subject, his conclusions are known today as Archimedes' Principle.

Perhaps you have heard or read the story of how Archimedes devised a way of telling whether the king's crown was made of pure gold. It would be well to read the story again in an encyclopedia as we study flotation and consider the importance of this information to our study of the structure of matter.

To study Archimedes' Principle as it is applied to floating bodies, you will need several wooden cylinders 5 to 10 cm long, cut from 2-cm ($\frac{3}{4}$-in.) doweling and varnished to make them waterproof; a 100-ml graduated cylinder; a pan or platform balance; water; and 50 ml of gasoline, kerosene, or alcohol. Half fill the 100-ml graduate with water, weigh it on the balance, and record the weight and the reading of the water level in the graduate. With the graduate still on the balance, carefully lower one of the wooden cylinders into the water. From the new reading of the water level you can find the amount of water displaced by the wood. From the new reading of the weight you can determine the mass of the wooden cylinder.

Divide the mass of the wooden cylinder by the volume of water displaced. Compare this number with the value of the density of water (1 g/ml). Suppose, instead of floating the wooden cylinder, you had added water to the graduate to bring the level of water to the same point.





Which liquid is denser?
(The two wooden blocks are the same size and have the same weight.)

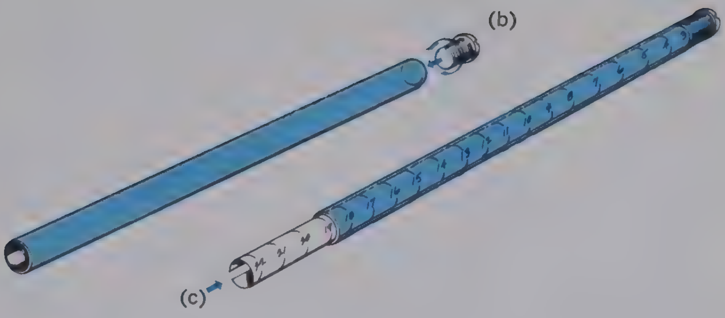
$\boxed{21}$ By how much do you believe this would have increased the weight of the water? After you have made a prediction and recorded it, check it by carrying out the necessary operations.

Repeat the experiment using wooden cylinders of different lengths that will fit in the graduate, or exchange readings with others who have used different samples. $\boxed{22}$ From these data try to draw a general conclusion that relates the weight of the solid samples and the weight of the water displaced during flotation, in the way that Archimedes did. $\boxed{23}$ What do you predict would be the relation between weight and water volume displaced if the specimen had been so dense that it would not float in water?

Instead of water, use 50 ml of gasoline or alcohol in the graduate and repeat all the operations and measurements. Although the change of weight should be the same when the cylinder is inserted, the change of volume of this liquid should be somewhat different. The volume of gasoline that must be displaced to have the same weight as the cylinder is not the same as the volume of displaced water because the density of gasoline is not the same as the density of water. You may also notice that the portion of the cylinder that sinks below the surface is not the same for the two liquids. $\boxed{24}$ Does the cylinder sink deeper into gasoline or into water? $\boxed{25}$ Which is denser, the gasoline or the water? $\boxed{26}$ Does this difference mean anything in relation to the structure of the substance?

Divide the volume of water displaced (the submerged portion of the cylinder) by the volume of gasoline that was displaced by the same specimen. The number you get is the relative density of gasoline com-

(a)



(a) Materials for constructing and using the soda-straw hydrometer.
(b) Putting the bolt in the straw.
(c) Inserting the scale.

pared to water, or the specific gravity of the gasoline. This wooden cylinder can be used to determine the specific gravity of any liquid in which it floats if the volume of the submerged portion of the cylinder can be measured. Used in this way, the cylinder serves as a hydrometer. A *hydrometer* is a floating instrument for determining the density of a liquid relative to the density of water by comparing the depths to which it sinks in the two liquids.

2 6 It is easy to build a hydrometer which is better suited than a wooden cylinder to most of the measurements you will make. You will need a transparent plastic drinking straw, household cement, lined paper, and a bolt that will fit in the straw. Cement the shank of the bolt firmly into one end of the straw. Cut a strip 4 mm wide from the side of a piece of lined paper; make it the same length as the straw plus the bolt head. Starting with the bottom of the paper as zero, number the lines in succession all the way to the top. Now cut the lower 5 cm from the paper strip, gently curve it lengthwise with the numerals on the outside of the resulting curve, and slide it into the straw until the top of the strip is even with the top of the straw. Place a drop of cement in the top of the straw to seal it and to hold the strip in place.

2 7 For using your hydrometer to obtain specific gravities, you will need a tall cylinder or test tube and the gasoline used in Experiment 2–5. Float the hydrometer in the cylinder of water. Read and record the depth to which the hydrometer sinks. Now place the hydrometer in gasoline in a tall container and record the depth to which it sinks. To

find the specific gravity of gasoline, divide the number that corresponds to the surface level when floating in water by the number in gasoline. We may write this as an equation:

$$\text{Specific gravity} = \frac{\text{level of hydrometer in water}}{\text{level of hydrometer in liquid}}.$$

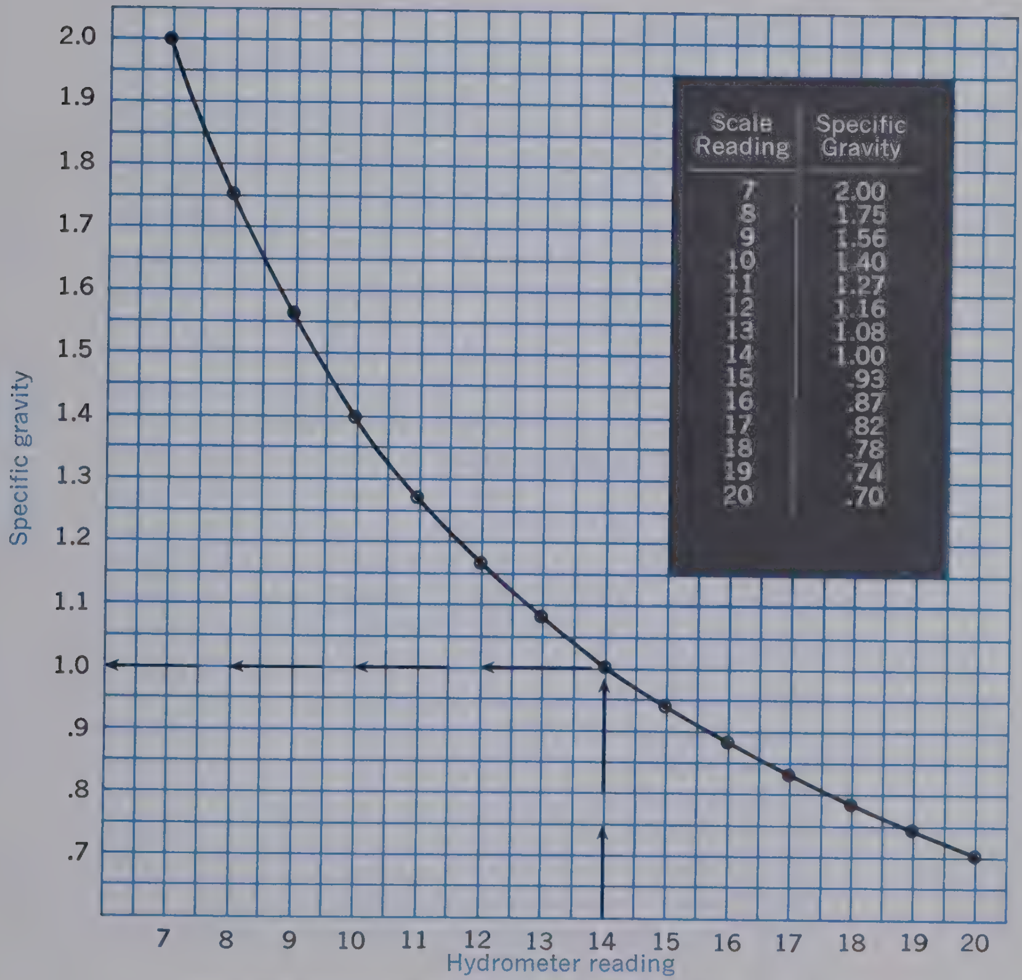
Check your results with those you obtained by using the wooden cylinder. Certain liquids, such as alcohol and carbon disulfide, may soften the cement or weaken the plastic straw if the hydrometer is left in them more than a few seconds. Be sure to rinse the hydrometer in water after immersion in other liquids.

Commercial hydrometers have a scale from which specific gravity may be read directly. You can indirectly read specific gravity with your hydrometer by preparing a graph in the following way. First divide each scale number into the number at the water level and record your results in a table. Then use specific gravity as the vertical scale and hydrometer readings as the horizontal scale on your graph, as shown in the sample. Record the data from your table as points on the graph, and join all the points. Now use your graph by getting the hydrometer reading for a liquid. Locate it on the horizontal scale, then go straight up from that point to the curve, then horizontally over to the specific gravity scale. You will find your hydrometer and graph useful in the study of solutions in Chapter 3. [F, G]

Specific Gravity of Solids by Flotation

In trying to determine the relative densities of the forms of sulfur, you may not have had very clear-cut results. You could not measure the volume of the samples directly because of their irregular shapes, so you measured the volume indirectly by displacement of water and then computed the densities. Even then, you may not have found a difference.

We can use Archimedes' Principle as a basis for another method of determining the relative densities of the forms of sulfur. Archimedes' Principle states that an object will float if it can replace a mass of fluid equal to its own weight and will sink if it cannot. If the densities of the object and the liquid are the same, the object will just barely float. Thus, if we can adjust the density of the liquid, we can determine the density of the object. To do this, we will use two liquids—one much denser than the forms of sulfur and one less dense. We will add one to the other to vary the density of the mixture.



Scale Reading	Specific Gravity
7	2.00
8	1.75
9	1.56
10	1.40
11	1.27
12	1.16
13	1.08
14	1.00
15	.93
16	.87
17	.82
18	.78
19	.74
20	.70

Can you say which is the densest form of sulfur?

2 8 You will need about 75 ml of tetrabromoethane (a liquid with a density of 2.96 g/ml); about 50 ml of butyl carbitol (a liquid with a density of 0.95 g/ml); fairly sizable samples of roll sulfur, rhombic sulfur, monoclinic sulfur, and amorphous sulfur, and an old piece of amorphous sulfur; two long-stemmed eyedroppers; your hydrometer; and several extra test tubes equipped with stoppers.

Put two of the sulfur samples in one of the test tubes and then add tetrabromoethane until the test tube is one-third full. Carefully add butyl carbitol with an eyedropper. Stopper and shake the tube between additions. [27] Can you figure out a way of predicting how much butyl carbitol has to be added before the sulfur will sink? Whether you can or not, add butyl carbitol until one or both samples sink. If both samples sink, try to refloat one by adding tetrabromoethane, drop by drop. [28] If you succeed in making one sample of sulfur float while the other still sinks, what can you say about the densities of the two samples? Stopper your test tube and save it for later use.

Use the same method to compare the densities of all the samples of sulfur. Finally, use your hydrometer to determine the specific gravity of the mixtures of tetrabromoethane and butyl carbitol that you have made for each test. Also weigh 20 ml of each of the solutions and calculate their densities.

As we go on with our investigation of the particle theory and the continuous theory, it will become quite apparent that one of the important properties of matter is density. The mass per unit volume is unique for each kind of matter. We even find that it varies for a single kind of

matter such as sulfur. We can interpret this last point to mean that the density of sulfur changes when the structure changes. As we study more samples of matter of many different kinds, we will try to determine whether the differences in density are due to different structural arrangements alone, or whether there are other reasons.

Review the evidence you have gathered from your study of density and specific gravity. [29] Does this information about the packing of matter cause you to favor the particulate theory of matter, the continuous theory, or do you feel that more investigations are required before drawing definite conclusions? [H, I, J, K]

TAKING INVENTORY

Curiosity has always led man to think about the world around him and to investigate it.

This chapter introduced you to past and present thoughts about the nature of matter. Two ideas have persisted through the ages. One is the view that matter is continuous; the other is that matter is particulate. A compromise view is that it is both.

This chapter also introduced you to some techniques of investigation which enable you to gather information about matter for yourself and make your own judgments about its nature. As part of your effort to acquire techniques, both laboratory and mental, you studied the forms of matter, you built models of crystals, you studied the relation of mass and volume, and you compared densities by determining specific gravity. As you go on to expand your investigation of the nature of matter, keep these basic techniques in mind and attempt to apply them. Chapter 3 concentrates on the form of matter known as solutions. What can a study of these mixtures of substances tell about the nature of matter?

FURTHER EXPLORATIONS

A. In addition to noting the changing fluidity of sulfur as it warms, you may want to examine the sulfur that condenses from the vapor near the top of the heated tube. Does it ever seem to exist as a liquid there, or does it suddenly change to the solid state? What is the nature of the solid sulfur you find there? Is it amorphous or crystalline? How would it look under a microscope? If you observe crystals, do they seem to be rhombic, monoclinic, or some other shape?

B. In the work you did with sulfur you poured liquid sulfur on a piece of filter paper in a funnel. The sulfur did not pass through the paper. Is this fact significant with regard to the theories of matter? What would happen if you poured water on a piece of filter paper in a funnel? Does sulfur appear to pass through the paper if it is dissolved in carbon disulfide?

C. The crystal pattern in ice can be demonstrated by a process known as etching. Put a sheet of fine sand paper on a table and rub one face of a hard-frozen ice cube across it to make the ice as smooth as possible. Return the ice cube to the ice compartment and refreeze it. Then rub this ice cube over a piece of soft cloth to melt a thin film of water on the cube and thus make the surface smoother. Again return the cube to the freezer. As the thin film of water freezes, it will build on the invisible crystalline structure of the cube, revealing a pattern. Examine the pattern for further insight into the structure of water.

D. What meaning can you attach to the fact that solid water floats in liquid water? What does this imply about the packing of particles in ice?

E. How can variation in density from one kind of matter to another be explained by the particle theory of matter? How can it be explained by the continuous theory of matter?

F. If a hydrometer sinks to a scale reading of 16.2 in water and to a scale reading of 13.8 in another liquid, what is the specific gravity of the second liquid?

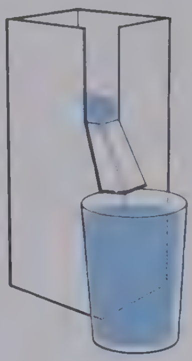
G. A piece of wooden doweling 1 cm in diameter and 20 cm long has a mass of 11.6 g. How much water will it displace when it floats? If the same piece of wood is immersed in kerosene which has a density of 0.82 g/ml, what weight of kerosene will it displace? what volume?

H. The density of any particular substance does not depend on the size of the sample tested. Since it is an expression of the ratio of mass to volume, changing the size of the sample changes these two quantities proportionally. Obtain several samples of the same substance that differ only in size. Roll sulfur might be used, but not samples representing the different forms of sulfur. By one of the methods suggested for determining density or specific gravity, find the density of each sample. Prepare a graphic plot of the volume as a function of mass. Can the points be joined by a straight line? State a simple formula for the density of the substance.

I. Air is a fluid similar to water. We know that some gas-filled balloons rise in the air. The air displaced by such a balloon has greater mass than the balloon and the gas it contains. Is there a like tendency for cork or Styrofoam to "float" in air? Would a sample of crushed Styrofoam weigh the same as it did before it was crushed?

J. Take a piece of aluminum foil about 6 cm square. Fold it into the form of a small open box. Put it in a beaker of water and observe whether it floats. Now remove the box from the water and wad it into a small ball. Return it to the water and observe what happens. How do you explain what you have observed?

K. Archimedes is said to have discovered the principle that bears his name as a result of observing a change in the water level of his bath when he entered it. If the tub had been full to overflowing, a volume of water equal to the volume of his body submerged would have spilled out. Construct an overflow device by cutting an open-top milk carton in the following manner. Cut down two adjacent sides of the carton along the dotted lines shown in the diagram. Make the cuts about 1 cm from the edge and about 5 cm long. Now fold the cutout strip down to make a spout from which displaced water may flow into a beaker or graduate.



Place the carton on a level table and add water until some flows out of the spout. When the flow has stopped, immerse an object in the container of water and catch the overflow to measure its volume. If the mass of the object is known, its density may be determined from these measurements.

Archimedes was also interested in shipbuilding. Would this interest have had anything to do with the development of his principle?

FURTHER READING

BOYS, CHARLES V., *Soap Bubbles and the Forces Which Mold Them*. Doubleday & Company, Inc., Garden City, N.Y., 1959.
Three lectures given by the author in 1889–90. Book is easy reading, and it is both informative and entertaining. Suggests many experiments that you may do which are related to this chapter.

BRAGG, WILLIAM, *Concerning the Nature of Things*. Dover Publications, New York, 1948.
A series of six lectures delivered by the author. Lecture I discusses the nature of matter; Lecture V discusses the nature of ice and snow.

CONANT, JAMES B., general editor, *Harvard Case Histories in Experimental Science*. Harvard University Press, Cambridge, Mass., 1957.
A two-volume set composed of eight case histories. Case 4, The Atomic-Molecular Theory, contains material related to the nature of matter.

LAPP, RALPH E., and the Editors of Life, *Matter*. Time-Life Books, Time, Inc., New York, 1963.
This book, which is a picture essay as well as a printed text, contains many ways of looking at matter. It is related to the content of Chapter 2 as well as almost all the other chapters of *MATTER AND ENERGY*.

LYNDE, CARLETON J., *Science Experiences with Inexpensive Equipment*. D. Van Nostrand, Inc., Princeton, N.J., 1950.

———, and FLOYD LEIB, *Science Experiences with Home Equipment*. D. Van Nostrand Company, Inc., Princeton, N.J., 1949.

———, and FLOYD LEIB, *Science Experiences with Ten-Cent Store Equipment*. D. Van Nostrand, Inc., Princeton, N.J., 1950.

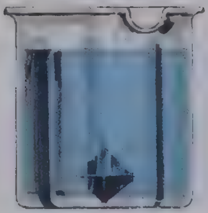
RUSSELL, BERTRAND, *The Analysis of Matter*. Dover Publications, New York, 1954.

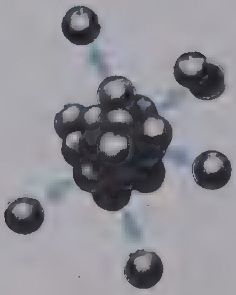
SCHWARTZ, GEORGE, and PHILIP W. BISHOP, editors, *Moments of Discovery*. Basic Books, Inc., New York, 1958.
This two-volume set contains a discussion and a reprinting of writings that represent the origins and the development of modern science.

3

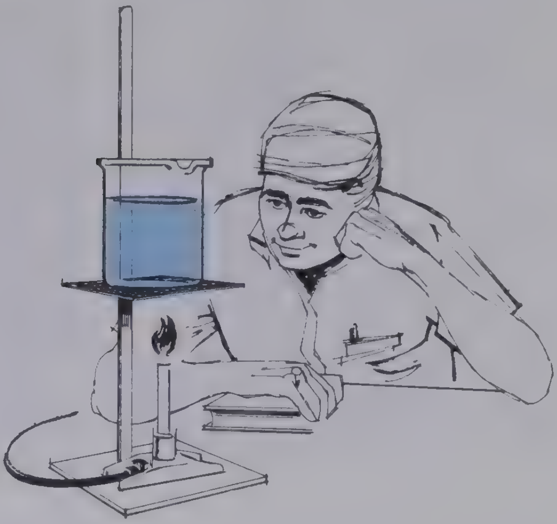
Learning More About Matter From Solutions

3-1 Characteristics of Solutions
3-2 Other Kinds of Mixtures
3-3 Uniformity in Preparing Solutions
3-4 Classification of Solutions
3-5 Equilibrium in Solutions
3-6 Temperature and Solubility of Solids
3-7 Temperature Changes That Accompany Dissolving
3-8 Theory and Explanation: The Function of a Model
3-9 Volume Changes in Solution





Salt may be separated from sea water by heating with a burner (right) or by the sun's rays in large ponds (opposite page). With both methods the water evaporates and salt deposits remain.

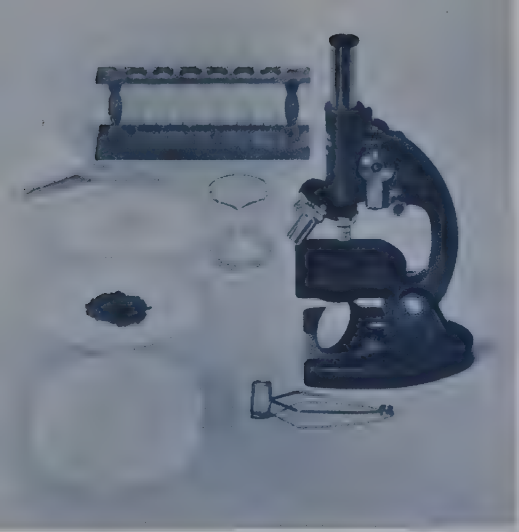


3–1 CHARACTERISTICS OF SOLUTIONS

You are familiar with soft drinks, coffee, tea, and lemonade. These are mixtures called solutions. Just what is a solution and how does this type of mixture differ from all others? In exploring our theories for the structure of matter, we have found evidence that matter may be particulate. A careful study of solutions will yield interesting information about matter and will help us to understand it better.

The chemist defines a *mixture* as an intermingling of two or more different kinds of matter which do not lose their identities. *Solutions* are mixtures with particular characteristics. There are many classes of solutions; one of the most common classes consists of a solid dissolved in a liquid, usually water. We refer to the solid as the *solute* and to the liquid as the *solvent*. You made such a solution in Chapter 2 when you dissolved sulfur in the solvent, carbon disulfide. By the physical process of evaporation of the solvent you were able to recover the solid sulfur. [1] What happened to the sulfur during the dissolving process? [2] Did the solution contain evidence of the sulfur? [3] Why was it possible to recover the sulfur in the form of rhombic crystals after the solvent evaporated? Let's prepare water solutions of sugar, sodium chloride (table salt), and copper sulfate, and attempt to determine the characteristics by which solutions can be identified.

You will need sugar cubes, rock (ice cream) salt, copper sulfate crystals, filter paper, a funnel, water, glass slides, a stirring rod, a 150-ml 3 1



Slowly heat the slide to dry the sugar solution.

beaker, six 20-by-150-mm test tubes, a test tube rack, a microscope, a source of heat, and your laboratory record book. You may find that you can organize your observations better by recording them in chart form.

Add water to the beaker until it is half full. Drop in one sugar cube and watch carefully as the sugar dissolves. 4 How can you tell when dissolving has stopped?

Taste the solution by dipping the stirring rod in it and touching the rod to your tongue. Now put a drop of the solution on a clean glass slide and see if crystals of sugar are visible when you view the mixture through a microscope. (If the slide has dust or lint on it, your observations may be misleading.) Put another drop of sugar solution on a clean glass slide and warm it *slowly* over a low flame to evaporate the water. When the water has evaporated, examine the slide for evidence of sugar crystals. 5 If present, have the sugar crystals been changed by the *dissolving process*?

Now let's see if the sugar can be removed from a solution by filtering. Fold a piece of filter paper as you did when working with sulfur, and place it in the funnel. Put the funnel in a test tube and pour about half the sugar solution on the filter paper. The liquid that passes through the filter paper and into the test tube is called the *filtrate*. Any solid that remains on the paper is known as the *residue*. Examine the filter paper for sugar crystals. Taste the filtrate and decide whether the sugar was removed by the filtering process. You may want to examine the filter paper under a microscope to see if there are visible holes through which the dissolved sugar might pass. 6 What can you say about the size of any possible particles of dissolved sugar?

Does filtering help to distinguish solutions, dispersions, and suspensions?

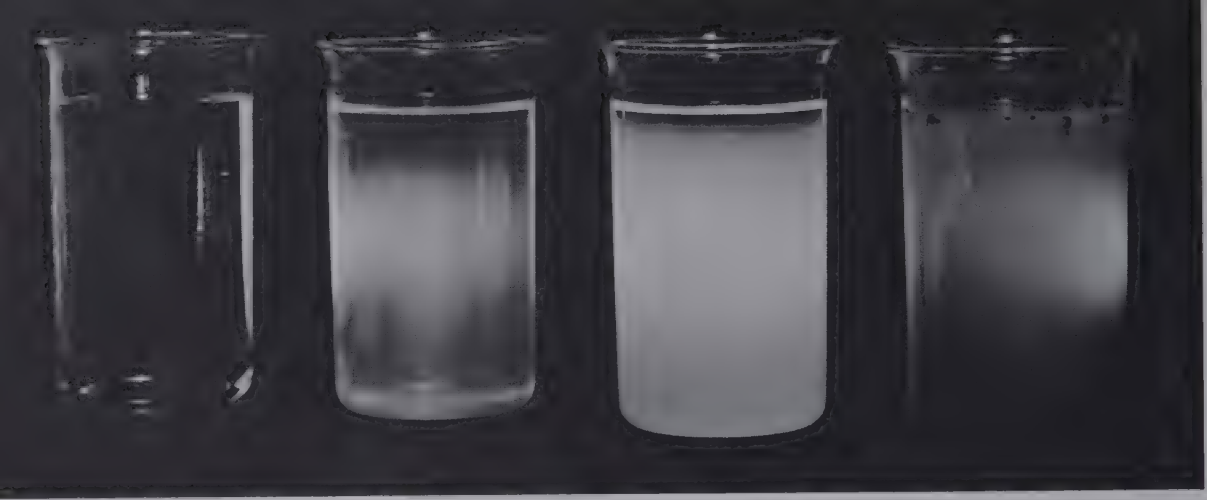
Repeat this series of tests, using three or four crystals each of rock salt and copper sulfate in about 50 ml of water in a beaker. **(CAUTION:** *Do not* taste the copper sulfate solution because *it is poisonous.*) These solids will dissolve much more readily if you crush the crystals before putting them in the water. When you have completed your observations on these two solutions, put them in separate test tubes, label and stopper them, and set them aside for a later experiment.

[7] What are the characteristics of the solutions you have worked with in these experiments? [8] Do your observations of these solutions lend support to the particle theory of matter? [A]

3–2 OTHER KINDS OF MIXTURES

Not all mixtures of solids in water are solutions. You may have noticed that when table salt is dissolved in water, the resulting mixture is not very clear. This is because certain materials of very limited solubility have been added to the salt to prevent caking in damp weather. Cloudy liquid mixtures like this are either dispersions or suspensions. Let's compare suspensions, dispersions, and solutions to see how they differ. We will also attempt to learn something about the size of the particles in these mixtures by filtering these mixtures and passing a light beam through them.

You will need the solutions you saved from the Experiment 3–1, six 20-by-150-mm test tubes, water, table salt, liquid laundry starch, finely crushed soil, a 100-ml graduated cylinder, a stirring rod, labels, a tea-



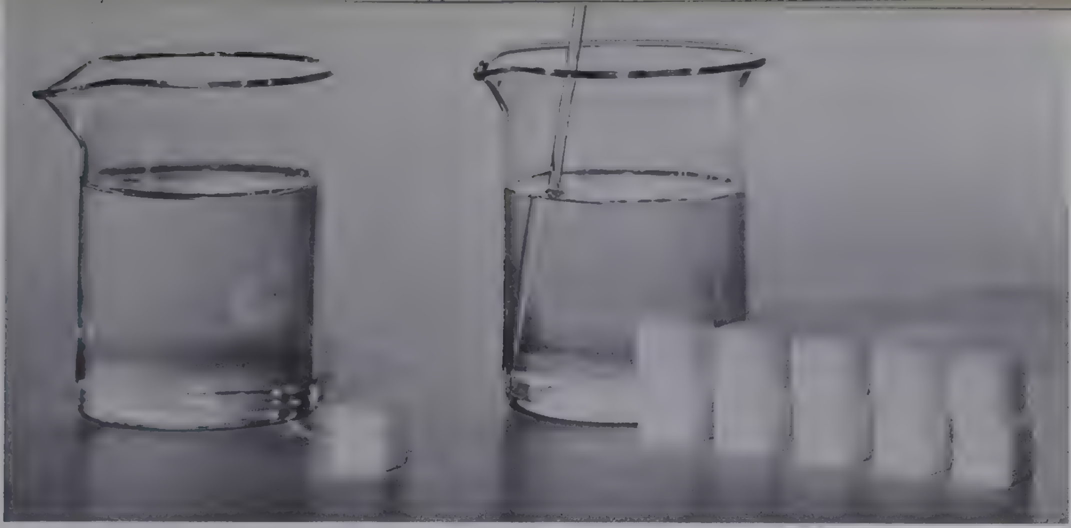
Which of these mixtures is a solution?

spoon, filter paper, a funnel, a test tube rack, a flashlight or projector, and your laboratory record book. Again, the results of this experiment will be more meaningful if you record them in chart form so that the characteristics of solutions, suspensions, and dispersions can be easily compared.

Put 15 ml of water in each of three test tubes. To the first add one teaspoon of table salt; to the second add one teaspoon of liquid laundry starch; to the third add one teaspoon of finely crushed soil. Stir each of the mixtures until no further change seems to occur. Filter about 5 ml of each mixture through separate pieces of filter paper, catching the filtrates in the three remaining test tubes. Label these test tubes and record any evidence of whether salt, laundry starch, or soil was removed by filtering. Set the original, unfiltered mixtures aside for several hours in order to determine which materials settle to the bottom of the test tube and which appear to remain uniformly distributed.

Shake each of the unfiltered mixtures so that the contents become thoroughly mixed again. Take these, along with your previously prepared solutions, into a darkened room. Place each test tube in a beam of bright light from a flashlight or a projector. Results will be more satisfactory if the light beam is made smaller by passing it through a hole about 2 mm in diameter before it strikes the samples. Pay careful attention to the appearance of the beam of light as it passes through the mixtures.

The mixture that transmits the light beam without affecting it is a solution. Notice the rather definite beam in the starch mixture. This reflection of light so that its path is clearly seen is called the *Tyndall effect*.



The beaker on the left contains a dilute sugar solution. The beaker on the right contains a concentrated sugar solution.

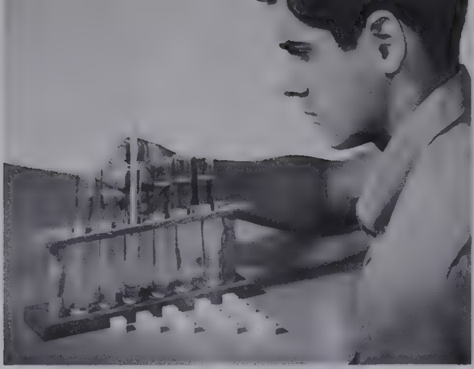
You have probably observed it in a movie theater when dust or smoke particles reflected the light from the projector. The same effect is evident on a foggy night when water particles in the air reflect light from the headlights of an automobile, or when a searchlight beam sweeps the sky.

Although the distinction between dispersions and suspensions is not definite, a *suspension* is usually classified as a mixture that exhibits the Tyndall effect and has particles which are large enough to be seen with the naked eye and settle out fairly rapidly to the bottom of the container.

A *dispersion*, on the other hand, is usually classified as a mixture that exhibits the Tyndall effect, but its particles are too small to be seen with a microscope, cannot be removed by filtering, and do not settle out after standing for a long time. Dispersions have a cloudy, nontransparent appearance. 9 How would you classify the starch and soil mixtures? 10 Do your observations of these mixtures lend support to the theory that matter is particulate? [B]

3–3 UNIFORMITY IN PREPARING SOLUTIONS

If you dissolve one sugar cube or a pinch of granulated sugar in 250 ml of water, the solution does not taste very sweet. This is a *dilute* solution. If you dissolve 25 sugar cubes in 250 ml of water, the solution tastes much sweeter. This is a *concentrated* solution, compared to the solution of one cube in 250 ml. *Dilute* and *concentrated* are relative terms. If each member of the class were asked to prepare a concentrated solution of sugar in



How does the specific gravity of sugar solution change as it goes from a dilute to a concentrated solution?

water and the results were compared, the amounts of sugar used by each student would probably vary widely, since each person would use his own opinion of how much sugar to dissolve. Therefore, when we use the term *dilute*, we simply mean that the solution contains much less solute than the solvent is capable of dissolving under existing conditions. The term *concentrated* means that the solution contains more nearly the maximum amount of solute that the solvent is capable of dissolving under existing conditions.

The hydrometer you constructed in Chapter 2 will help you distinguish between the very dilute and the more concentrated solutions of some solutes. You may also gain some insight into the nature of solutions and the structure of matter from a study of the relationship between the concentration of solute and the specific gravity of the mixture.

You will need a 100-ml graduated cylinder, a stirring rod, your drinking-straw hydrometer, water, several sugar cubes, and the specific-gravity graph you made in Experiment 2–5. Measure 10 to 15 ml of water into the test tube so that it is deep enough to float your hydrometer. Add one sugar cube, dissolve it by stirring, and obtain a reading of the specific gravity by reading the hydrometer level and consulting your specific-gravity graph. Record these data and begin constructing a graph showing the specific gravity as a function of the number of cubes dissolved.

Add more sugar, one cube at a time, repeating your observations and recordings until no more sugar will dissolve.

If we are to have uniformity in the preparation of solutions, we need more precise methods for expressing the concentration of solute and solvent. One such method is to use *percent by mass* (weight) of solute. According to this method, if you want a 5 percent sugar solution, you add 5 g of sugar to 95 g of water. If you dissolve 2 g of salt in 98 g of water, you have a 2 percent salt solution. Since 1 ml of water weighs 1 g, you can simply measure the amount of water required for each solution in a graduated cylinder and save the time of weighing it. However, if you want to use another solvent, such as alcohol or carbon disulfide, then it must be weighed, as 1 ml of these solvents does not weigh 1 g.

If we knew the mass of each sugar cube, we could calculate the percent by mass of sugar in the solution for each of the concentrations in the preceding activity. We could weigh one sugar cube, and from this information, calculate these concentrations. However, since the possibility of error is high for this method, we can partially avoid error by a method of averages. We believe that each sugar cube has about the same mass; therefore, if we weigh ten and divide this mass by the factor ten, we will arrive at a value for the average mass of a cube that is probably more nearly the correct value than a figure obtained by one individual weighing. If the cubes have been used in previous experiments, they may be worn and thus quite variable in mass. If they are fresh cubes and you have the box from which they were taken, the total mass in grams of the sugar in the package divided by the number of cubes will give you a way to check your weighings. Determine the average mass of a sugar cube using both methods just described, and calculate the percent by mass of sugar in the solution for each of the concentrations in the preceding activity. Add this information to the graph already drawn, or construct a new graph with specific gravity as a function of concentration.

Another method for preparing uniform solutions is by *percent by volume*. This is commonly used when both parts of the solution are liquids. For example, if you want to prepare a 10 percent solution of alcohol in water, you will use 10 ml of alcohol and add enough water to make 100 ml of solution. Such concentration of solutions is based on the volume of solute in the total volume of solution. (The chemist uses other methods for insuring uniformity in solutions, but we will not use them in this course.) [C, D, E]

3–4 CLASSIFICATION OF SOLUTIONS

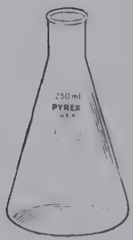
We may classify solutions as unsaturated, saturated, and supersaturated, depending on the amount of solute present in a given quantity of solvent under certain conditions. A solution is *unsaturated* if more solute can be dissolved in the solvent without changing the temperature. If no more solute can be dissolved in the solvent at that temperature, the solution is said to be *saturated*. If a solution is holding more solute than could normally be dissolved in that solvent under existing conditions, then it is said to be *supersaturated*. In Experiment 3–3, your sugar solution was unsaturated until more sugar was added than could be dissolved. At this point the solution became saturated.

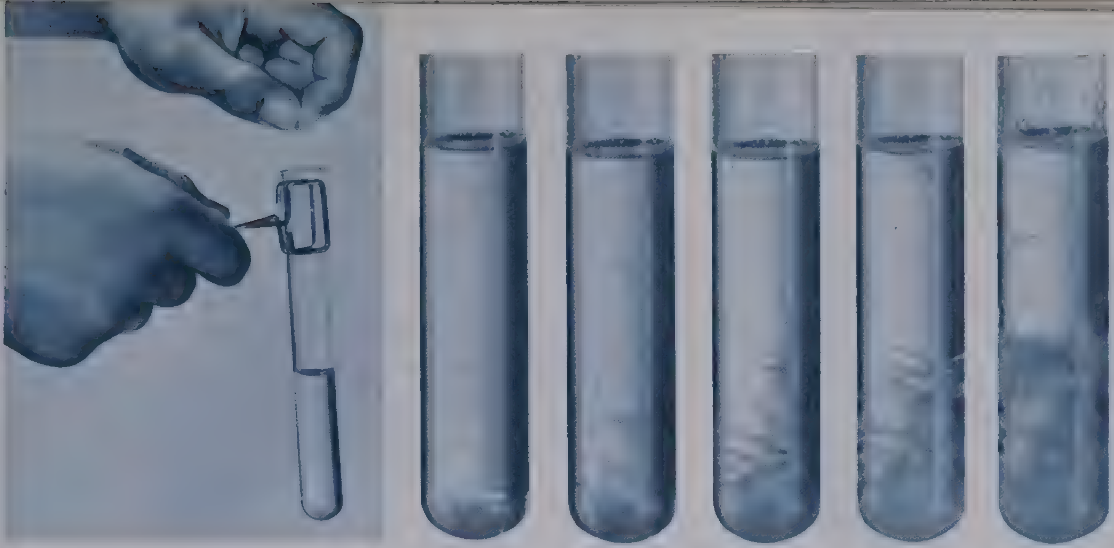
Weighing out the hypo for Experiment 3–4.

Although a supersaturated solution is difficult to prepare, you can obtain one. You will need a balance, a 20-by-150-mm test tube, a test tube rack, a stirring rod, some distilled water, a 100-ml graduated cylinder, a 250-ml beaker half full of cold tap water, a test tube holder, a source of heat, and sodium thiosulfate crystals, commonly called *hypo*.

Put 2 ml of distilled water in the test tube. Weigh out 15 g of hypo. Put four of the hypo crystals in the test tube of distilled water and stir. When the four crystals have dissolved, you have an *unsaturated* solution, because more hypo can still be dissolved in this amount of water without changing the temperature. Continue to add hypo crystals, one at a time. Stir the solution to aid the dissolving process. When no more crystals will dissolve in the water at room temperature, you have a *saturated* solution. Note the small amount of hypo crystals you have used up to this point.

Now gently heat the solution, adding hypo crystals until all have been dissolved. (Be careful *not* to allow the solution to boil as this would reduce the amount of water.) When all the hypo has been dissolved you have a hot, concentrated solution. Place the test tube of solution in the beaker of cold water and allow it to cool to room temperature without disturbing it. We know that this solution will reach a point of saturation at some temperature during the cooling process, because only a few grams of hypo could be dissolved in water at room temperature. Observe whether any crystals settle out as the solution cools. If they do not, you have successfully prepared a *supersaturated* solution. You noted that a





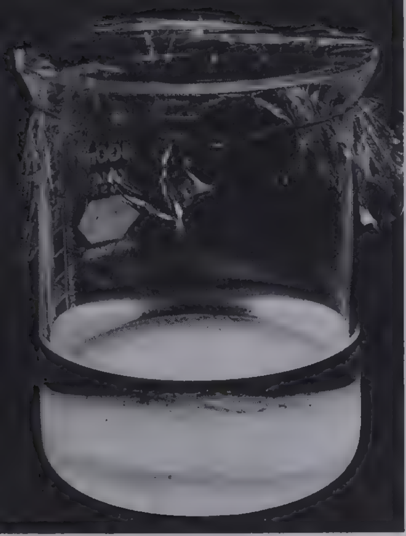
What changes do you notice in the stages from the supersaturation to the crystalline stage?

very small amount of the 15 g of hypo would dissolve in the water at room temperature. Yet now the entire 15 g of hypo is in solution at room temperature. This conforms to our definition of a supersaturated solution: one which holds *more* solute than could normally be dissolved at the existing temperature.

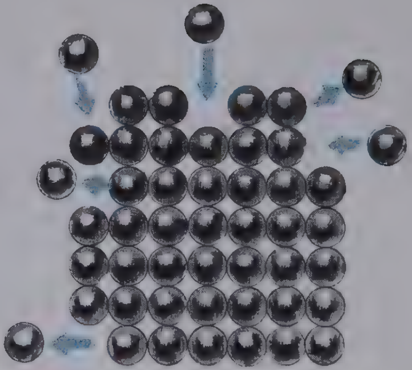
You are now ready to observe a most interesting reaction. Carefully lift the test tube of supersaturated solution from the beaker of water. Note the approximate temperature of the solution by touching the lower portion of the test tube to your inner wrist. While you watch closely, drop one crystal of hypo into the solution. [11] What happens to the solution? As the reaction progresses, note the temperature of the mixture from time to time by touching the test tube to your inner wrist. [12] Can you explain this temperature change? [F, G, H]

3–5 EQUILIBRIUM IN SOLUTIONS

As you continued to add hypo to water, there came a point at which no more would dissolve at room temperature. The crystal remained at the bottom of the test tube and was apparently not changing. The solution was saturated. But has the dissolving process stopped? To answer this question we will use a saturated solution of another substance and crystals of that substance.



Could equilibrium exist at the crystal surface if the beaker were open?

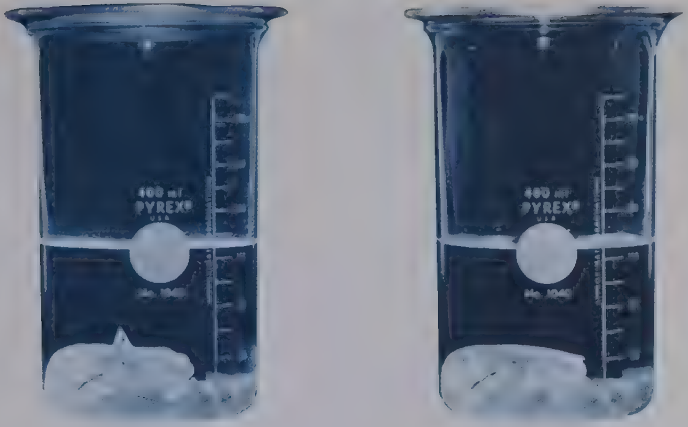


When a solution is saturated the rate of dissolving and the rate of crystallization are equal.

3 5 You will need 10 g of alum (potassium aluminum sulfate), a 250-ml beaker, a stirring rod, a 100-ml graduated cylinder, a rubber band, and an 8-cm square of plastic wrap. Measure out 50 ml of water and pour it into the beaker. Add the 10 g of alum and stir the solution thoroughly. Cover the beaker with the square of plastic wrap and secure it with the rubber band. Set this solution in a quiet place where the temperature is fairly constant. After 24 hours, examine it. Remove the cover and use the stirring rod to break up the crystals in the bottom of the beaker. Cover the beaker again and then make a rough sketch of the appearance of the crystals for a later comparison. Then carefully put the beaker away. After another 24 hours, examine it again. [13] Do the crystals look the same as they did the day before? [14] Can you detect any evidence of change in the arrangement of the crystals? [15] Is there evidence that indicates that dissolving and crystallizing are both taking place?

This experiment and others like it lead us to conclude that hypo and other substances continue to dissolve after the point of saturation is reached. The dissolving only *appears* to stop because an opposite process of crystallization was going on all along and has caught up with the rate of dissolving when the solution is saturated. When the rate of dissolving equals the rate of crystallizing, a condition called *dynamic* (active) *equilibrium* exists.

The more solute there is in solution, the greater the probability that some of it will come into contact with the solid at the bottom of the test tube and again take its place in the crystal. As the hypo in solution rejoins



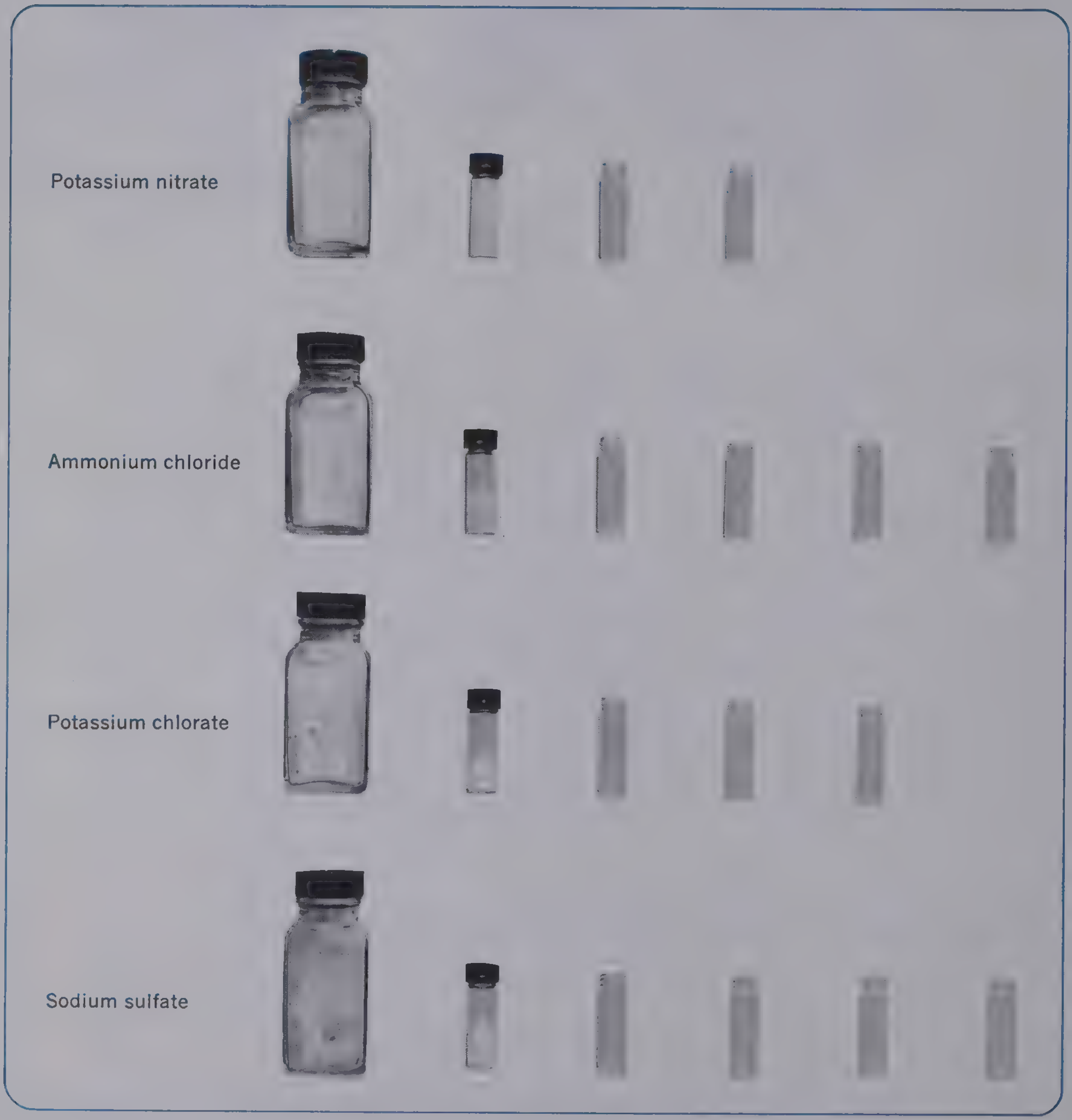
The change in the contour of the crystal surface over a period of four days shows that there is an exchange between the solution and the crystals.

the crystal, other particles of hypo leave the solid form and go into solution. A kind of substitution process is going on much as in a game of football. According to the rules, the number of players on the field must remain constant regardless of how many boys are on the bench. In order for a player from the bench to get into the game, one of the players on the field must come out. So it is with the saturated solution. As dissolved hypo rejoins the crystal, other hypo leaves the solid form and goes into solution. In a saturated solution, crystals are constantly in the process of dissolving and crystallizing. Thus, the solid at the bottom of the container is in equilibrium with the solute dissolved in the liquid. Whether there is only one undissolved crystal or enough to cover the bottom of the container is not important. The rate at which the solid is going into solution exactly equals the rate at which dissolved solute is again becoming a part of the solid. Similar dynamic balance is found in many natural processes, and we will encounter other examples of such equilibrium as we continue our studies. [I]

3–6 TEMPERATURE AND SOLUBILITY OF SOLIDS

In order to dissolve 15 g of hypo in 2 ml of water, it is necessary to heat the mixture. $\boxed{16}$ Does increasing the temperature of the solvent always make it possible to dissolve a greater amount of solute? $\boxed{17}$ Do solids vary in the amount that can be dissolved in a given amount of

Label your sample containers carefully for Experiment 3-6.



Be sure to record the weight on the labels as well as in your laboratory record book.

water at any given temperature? We will study four different solids in an attempt to answer these questions.

You will need three 250-ml beakers, a source of heat, a 20-by-150-mm test tube, a 10-ml graduated cylinder, a test tube holder, a ring stand, a ring, an asbestos pad, a test tube rack, a thermometer, a one-hole stopper to fit the test tube, distilled water, ice, potassium nitrate, ammonium chloride, potassium chlorate, sodium sulfate, a balance, a mortar and pestle, 17 small bottles or vials with caps for storage of solid samples after weighing, and 17 labels.

Weigh the following samples and store them in *stoppered* marked bottles so they will be ready as you need them.

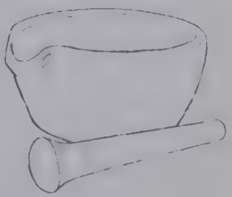
Potassium nitrate—Weigh out *three* 5-g samples. These will be added to the water one at a time.

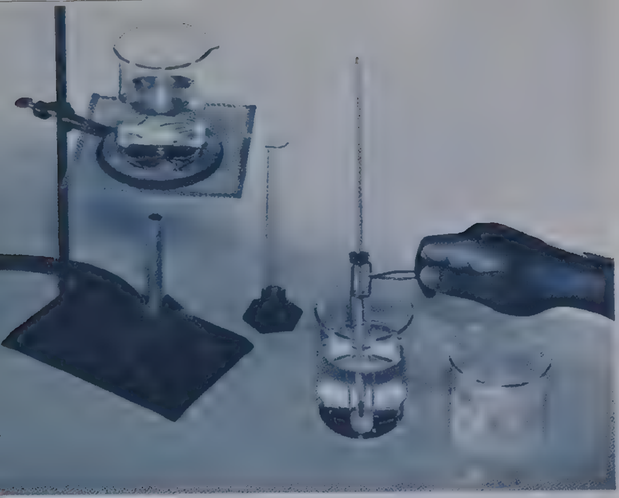
Ammonium chloride—Weigh out *five* 1-g samples to be used one at a time.

Potassium chlorate—Weigh out *four* 1-g samples to be used one at a time.

Sodium sulfate—Weigh out *five* 2-g samples to be used one at a time.

In this experiment we will try to determine the carefully measured temperature at which each sample of solute dissolves. To do this you must insert the thermometer into the test tube stopper. (**CAUTION:** Be sure to follow your teacher's instructions in carrying out this operation.) Moisten the stopper and thermometer and carefully insert the thermometer into the stopper hole.





Place the water baths close enough together so that the test tube can be transferred easily from one bath to another.

Use the three 500-ml beakers to set up three water baths, as shown in the illustration. The first should be an ice-salt bath to bring the test tube of water to 0°C. The second should be water at room temperature and the third should be water near or at the boiling point.

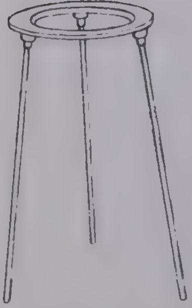
Carefully measure out 10 ml of distilled water. Pour the water into the test tube and cover the test tube loosely to avoid excess evaporation. Put the test tube into the ice-salt bath and cool it to 0°C. Add one of the solute samples and shake the solution gently as it dissolves to insure uniform mixing; note as precisely as possible the temperature at which the sample disappears. You may find that the sample does not dissolve completely at temperatures close to 0°C. If not, transfer the test tube to the water bath at room temperature and allow it to warm until the sample does dissolve.

By transferring the test tube back and forth from one bath to another, you will be able to control the temperature quite precisely. Add a second sample to the test tube and determine the lowest temperature at which it remains dissolved. Use the hot water bath to bring the temperature to points above room temperature. Again, transfer the test tube back and forth between water baths to control the temperature.

Repeat the entire procedure with each of the remaining substances. Rinse the thermometer and test tube thoroughly before adding a new 10-ml quantity of distilled water. [18] Did you notice whether there was much evaporation from your test tubes? How would this affect the results?

Plot the mass of dissolved solute (vertically) as a function of the temperature (horizontally).

[19] What conclusions can you draw about the relationship between the solubility of these solids and temperature? [J]



Heating the copper sulfate crystals.

3–7 TEMPERATURE CHANGES THAT ACCOMPANY DISSOLVING

If you put an ice cube in a beaker of water, you will observe that it slowly disappears much like the substances you have just studied. The warmer the temperature of the water, the more rapidly the ice melts. Therefore, it would seem that melting and dissolving are similar processes. We also know that as ice melts it cools the water. [20] Do other solids take heat from water as they dissolve?

3 7

To investigate this question you will need a thermometer, six 20-by-150-mm test tubes, water, a teaspoon, the crystal model built with marbles and wax in Chapter 2, a 250-ml beaker, a mortar and pestle, a ring stand, a ring, an asbestos pad, a source of heat, ammonium chloride crystals, potassium nitrate crystals, sodium sulfate crystals, copper sulfate crystals, a test tube holder, a test tube rack, and a 10-ml graduated cylinder.

Crush all the crystals into small particles using a mortar and pestle. Put 10 ml of water in each of four test tubes. Note the temperature of the water in one test tube. Add $\frac{1}{2}$ teaspoon of ammonium chloride crystals to the water. When the crystals have dissolved, read the temperature and record whether it has changed or not. If it did change, record how much.

Repeat this procedure three times using the potassium nitrate crystals, the sodium sulfate crystals, and the copper sulfate crystals in place of the ammonium chloride crystals. Be sure to record the results in your laboratory record book.

Certain crystals, such as copper sulfate and hypo, contain water as a part of their structure. Put a single, large crystal of copper sulfate into

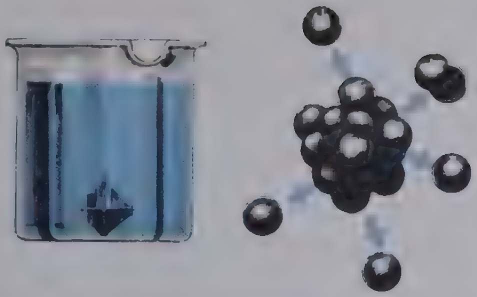
How will the heat affect the crystal model?

a *dry* test tube. Hold the test tube over a flame, as shown in the illustration, until all the water has been driven off. Note any color change that occurs. Carefully move the test tube over the flame until all moisture has evaporated from the inside of the tube. Empty the contents onto a piece of paper. This form of copper sulfate is known as *anhydrous* (without water) copper sulfate. [21] Do you find any evidence of crystals? Examine some of the material under a microscope. When you have finished studying this anhydrous solid, place $\frac{1}{2}$ teaspoon of it in 10 ml of water and measure the temperature change which it produces in the water just as you did previously with the four solids.

[22] Can you propose a hypothesis to explain why certain solids cause the water temperature to rise as they dissolve, while other solids cause water temperature to drop during the dissolving process? For clues, take the crystal model you built with wax and marbles and put it in a beaker of water using a ring stand and ring. Gently heat the beaker and watch what happens to the crystal model when the water becomes warm. We must remember, however, that the marble pyramid is *not* a crystal and the forces acting here are not the same as those that hold a crystal together. [23] Using this experience for comparison, what do you think may have been occurring as some of the solids dissolved in water? [K]

THEORY AND EXPLANATIONS:
3–8 THE FUNCTION OF A MODEL

A good scientific theory enables us to explain what we see. One theory to explain the structure of matter is that matter may be composed of very small particles, far too small to be seen even with a good microscope. In many ways the ice, ammonium chloride, potassium nitrate, and sodium



A model showing the dissolving of a crystal in water.

sulfate behave just like the crystal model when they are placed in water. If we assume these solids to be made of tiny particles, as was our model, is it not logical to assume that as these particles *absorb heat* from the water, they acquire enough energy to move apart from each other? In our model the marbles were held in an orderly arrangement by paraffin. They were not free to move around until the paraffin melted and released them. When this happened, the model fell apart and complete disorder was the result. Perhaps as these solids dissolve, disorder increases until the particles move so far apart that the crystals lose their structure. Of course, particles are not held to one another by something like wax. Our experimentation has not given us clues about what this binding agent might be, but that need not prevent our wondering about it.

3–9 VOLUME CHANGES IN SOLUTIONS

As you discovered in your experiment on temperature and the solubility of solids in liquids, the volume of the final solution was always greater than the original volume of the solvent. Does the final volume of the solution *always* equal the sum of the volumes of solute plus solvent? Measuring actual volumes of solid solute is difficult, because the solids we have used are in small, irregular pieces. It is almost impossible to determine the actual volume of sugar in a spoonful of grains or in one cube, since we do not know how much of the apparent volume is taken up by air spaces. (You saw evidence of air trapped in a sugar cube as you watched the cube dissolve.)

Since liquids do not contain trapped air, we can use a liquid solute to compare the volume of solution with the sum of volumes of two separate substances.

Does the combined volume of alcohol and water equal the volumes when they are apart?

3 8

You will need water, ethyl alcohol, a 100-ml graduated cylinder, and two smaller graduated cylinders which can be read to 0.1 ml.

Pour a carefully measured 27 ml of alcohol into the 100-ml graduated cylinder. In one of the small cylinders pour a carefully measured 3 ml of water. Then pour the 3 ml of water into the alcohol and stir. Note the final volume of this solution. Now pour 5 ml of water into the small cylinder you have used for water and 5 ml of alcohol into the other small cylinder. Add the 5 ml of alcohol and then the 5 ml of water to the 100-ml graduate. Stir again and note the volume. Continue adding the quantities of alcohol and water shown in Table 3–1 until you have all the combined volumes.

Table 3–1

Volume of alcohol added	Total volume of alcohol	Volume of water added	Total volume of water	Sum of separate volumes	Percentage of alcohol by original volume	Volume together
27 ml	27 ml	3 ml	3 ml	30 ml	90%	......
5 ml	32 ml	5 ml	8 ml	40 ml	80%	......
10 ml	42 ml	10 ml	18 ml	60 ml	70%	......
0 ml	42 ml	10 ml	28 ml	70 ml	60%	......
0 ml	42 ml	14 ml	42 ml	84 ml	50%	......

If you had added 3 ml of water to 27 ml of water instead of to 27 ml of alcohol, the final volume would have been 30 ml. [24] Why should the mixture of alcohol and water be different? Recall the crystal models you made with marbles. [25] Is it possible that spaces exist between the particles of water which can be occupied by alcohol particles or vice versa? We can make this analogy: When a bucket is filled with marbles, it can still hold a certain amount of sand without overflowing. [26] Does the experiment on volumes of alcohol-water mixtures lend support to the particle theory of matter?

TAKING INVENTORY

We have found that solutions are *mixtures* in which the solute and solvent are uniformly distributed; the solute and solvent seem to be composed of tiny particles; solute particles in solutions do not settle out upon standing; the particles in a solution are too small to reflect light (to give the Tyndall effect). Dispersions and suspensions are mixtures similar to solutions, but they have larger solute particles.

Solutions may be unsaturated, saturated, or supersaturated, depending upon the amount of solute dissolved in the solvent. Temperature and the nature of the solute determine how much solute can be dissolved. Supersaturated solutions are unstable. Saturated solutions are in a state of dynamic equilibrium with crystals of the solute substance.

Certain solutes lower the water temperature as they dissolve, while others raise the water temperature as they dissolve.

The principles learned in this study of solutions should enable you to produce crystals. The variety of shapes and colors found in crystals should greatly increase your understanding of matter. You have already produced sulfur crystals and have observed several crystalline salts in your study of solutions. The next chapter is devoted to an intensive study of the patterns of crystals.

FURTHER EXPLORATIONS

A. There are three states, or phases, of matter: solid, liquid, and gas. There are nine classes of solutions. In this chapter, we are working mainly with only one class—a solid dissolved in a liquid. Try to list the other classes and think of examples of them. Remember that to be classed as a solution, the mixture must exhibit the properties of solutions which you have identified.

B. Apply the terms *solution*, *dispersion*, and *suspension* to atmospheric conditions. For example, how would you classify fog or dust?

C. What mass of salt must you add to 100 ml of water in order to change the specific gravity enough so that it may be detected with your hydrometer? How many grams of salt must you add to 100 ml of water before you can detect the presence of the salt by taste? If you live in an area where there are bodies of salt water, judge the approximate concentration of salt in a sample of the water by determining the specific gravity.

D. Finding the average mass of several sugar cubes has been recommended as a better way to represent the mass of individual sugar cubes than weighing a single cube. To investigate the validity of this recommendation, make a great many individual weighings of sugar cubes and record them in a table. Do as many careful weighings as your patience allows.

After you have weighed the sugar cubes several times, determine an average mass for all of them. What is the greatest deviation from that average? Can you identify a range of deviation around the average that would include 60 to 70 percent of your weighings? Can you devise a means of graphing the data of your table?

E. Look for evidence of a nonuniform distribution of solute, as when a sugar cube is allowed to dissolve in water without stirring. Place one sugar cube in a test tube containing 25 ml of water. Observe the currents near the cube as it dissolves. Try suspending another cube near the top of the water in a second test tube. Is the density of this solution the same near the suspended cube as in other parts of the water?

F. What would you predict to be the density of a supersaturated hypo solution? Two g of water has a volume of 2 ml. The density of hypo is 1.73 g/cm^3. Fifteen grams of hypo has a volume of 8.7 ml. If there is no change in volume because of solution, we might expect the total volume to be 10.7 ml. The density, therefore, would be

$$\frac{17\text{ g}}{10.7\text{ ml}} = 1.6\ \frac{\text{g}}{\text{ml}}$$

Was this the volume of the solution you made? Was there any change in volume when the hypo crystallized from the supersaturated state?

G. Prepare a supersaturated solution of hypo by dissolving 15 g in 2 ml of water. When the solution is cool, drop a copper sulfate crystal (about 2 mm in diameter) into the solution. Compare the results with what you observed when a hypo crystal was added.

H. Put your knowledge of classification to work with your knowledge of solutions. How many different schemes for classifying solutions have been described? Try to work out a scheme of your own.

I. You can satisfy yourself that a saturated solution is a system in dynamic equilibrium. Grow a crystal of sulfur (or alum or copper sulfate) until it is large enough to weigh and measure accurately. Chip a corner off this crystal.

Record the weight and dimensions of the damaged crystal, then place it in a saturated solution of the same solute. Seal the container to prevent evaporation and set it aside. After several days, see if the damaged corner is repaired. Remove the crystal from solution, dry it, and again weigh it and measure its dimensions. Do your results indicate that some of the crystal returned to the solution at the same time that some of the dissolved solute was being deposited to repair the damaged corner?

J. Most canned or bottled soft drinks are called carbonated beverages because an abundance of carbon dioxide is dissolved in them. By observing the behavior of the dissolved gas after you have opened a cold bottle of such a beverage, try to determine the law or laws that describe the solution of gases in water. What classification system would you suggest for concentration of gases in liquid? How can a solution of a gas in water be made a supersaturated solution? How can you obtain evidence regarding the solutions of gases in tap water?

K. Repeat the drying process of a copper sulfate crystal from Experiment 3–7, but this time weigh the test tube and crystal before heating and weigh the test tube and powder after heating. What is the percentage loss of weight in the copper sulfate crystal when the water is removed? Does the sample recover this full amount when water is added again to the anhydrous copper sulfate?

FURTHER READING

Handbook of Chemistry and Physics. The Chemical Rubber Publishing Company, Cleveland, 1966–67.
A valuable source of tabulated experimental information. Chemists and physicists consider this book an invaluable tool to their work and study.

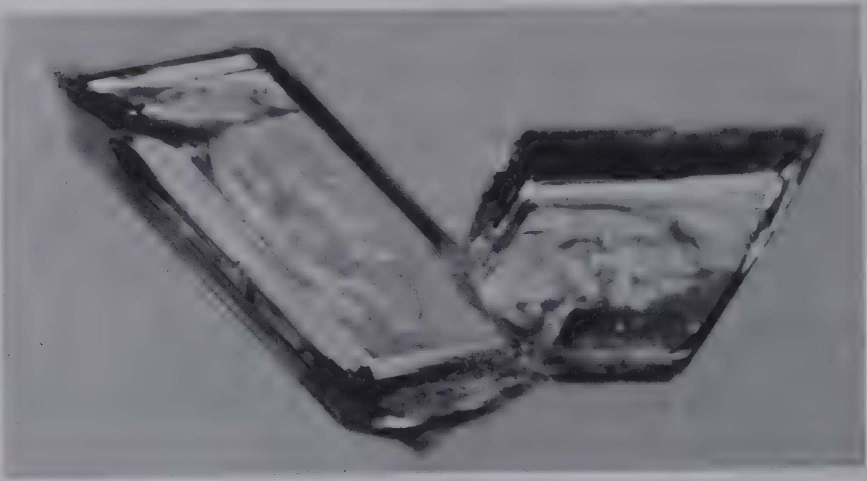
HELFMAN, ELIZABETH A., *Water for the World.* David McKay Co., Inc., New York, 1960.
This book discusses water as a major concern of modern man. The social implications as well as the scientific are considered.

HOLDEN, ALAN, and PHYLIS SINGER, *Crystals and Crystal Growing.* Anchor Books, Doubleday & Company, Inc., Garden City, N.Y., 1960.
Pages 58–92 contain a good discussion of the kinds and characteristics of solutions. Also contains many diagrams and solubility charts.

LAPP, RALPH E., and the Editors of Life, *Matter.* Time-Life Books, Time, Inc., New York, 1963, pp. 73–88.
Contains information on solutions along with excellent and attractive color pictures. Included are a color picture of the Tyndall effect, a rainbow, and familiar food mixtures to illustrate solutions and suspensions.

SEABORG, GLENN T., "Our Heritage of the Elements." *Chemistry*, Vol. 38, No. 5 (May, 1965), pp. 14–18.
A brief tracing of the discovery of the elements by a man who helped to discover nine of them.

4 The Patterns of Crystals



4-1 Some Characteristics of Crystals

4-2 Testing the Particle Theory

4-3 Other Tests of the Particle Theory

4-4 Metallic Crystals and the Particle Theory

4-5 Crystal Formation by Precipitation and Freezing

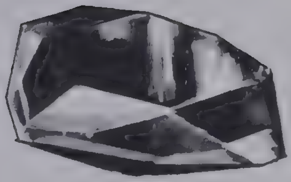
4-6 Crystal Patterns

4-7 Classification of Crystals

4-8 Crystal Patterns and the Nature of Matter

4-9 Variations in Crystal Patterns

4-10 Variety of Particles



4–1 SOME CHARACTERISTICS OF CRYSTALS

The impressions we get of something are determined by our perspective on it. The gentle burning of a candle and the holocaust of a forest fire are basically the same kind of event, but they seem different to the casual observer.

Our study of density, the changes we observed in sulfur, and our construction of marble models all strongly favor a particle theory of matter. However, we may not yet rule out all other theories, since the extent and variety of our observations are still quite limited. There are forms of matter we have not yet studied which may give us different viewpoints and force us to alter our total concept of matter.

In this chapter we will study more about solids; then in Chapter 6 we will turn our attention to gases. Until our picture is more complete, we must consider our conclusions about the structure of matter to be tentative.

As we began to look for the best model for the structure of matter, we examined crystals of water and sulfur for clues. Let's review the characteristics of crystals that we have observed.

1. Crystals form from gases when they cool, as when frost formed on the ice-filled can or sulfur crystallized in the neck of the test tube.
2. Crystals form from liquids as they cool. This occurred in the hypo solution. Remember that the rapid formation of hypo crystals caused a rise in temperature. As liquid sulfur crystallized, we did not attempt to detect any such change. This does not mean that it did not take place.
3. Crystals form from solutions when the solvent evaporates.
4. Crystals of any particular substance produced in the same way seem to have similar shape, color, and to a certain extent, size.

How does magnification affect your view of crystals?

4–2 TESTING THE PARTICLE THEORY

Although each of the characteristics of crystals mentioned above can be interpreted as behavior of particles, they might just as strongly support some other description of matter that we have not yet considered. To further test the particle theory, we might follow another scientific procedure in which we make predictions based on our particle theory and then test these predictions, one at a time, by experiment.

For example, we have observed what happens in a saturated solution when the dynamic equilibrium between dissolved and undissolved solids is disturbed. If some of the solvent evaporates from a saturated solution that is in contact with the solid solute, the crystal will grow from the addition of the solid that can no longer stay in solution. Can we use our particle theory to describe occurrences at the submicroscopic level that will explain what can be observed?

As part of your work in Chapter 2, you saw sulfur crystals form in a sulfur-carbon disulfide solution as the solution evaporated. Lowering the temperature also causes the amount of crystal solute to increase. [1] What might have happened at the particle level to produce the regular crystalline form that you observed? [2] Can the particle theory be used to explain what you observed and to predict what will happen if the conditions are changed?

Think about the large models we made with marbles. We cannot regard these as authentic large-scale models of crystals, but they can be helpful as we think about the possible arrangement of particles at the submicroscopic level. Like crystals, the marble models have plane surfaces with fixed angles between faces. We fashioned the models by carefully putting the marbles, one by one, in places where they seemed to fit. Imagine what would happen if your whole class began to build a

single crystal model with marbles and a cement that dries on contact. Let's say that each student must add two marbles to the model every minute. For a class of 30 students, the growth rate would be 60 marbles per minute, or one marble per second.

Our first problem would be deciding whose marble should be the first one in the stack. Once two marbles have been brought together, there is a natural place for the third to fit, and building can go on. But picture the confusion that might result if 30 people each tried to place two marbles in this tiny model each minute. Contact between some marbles might occur before they reached the central "crystal," and new crystals would start "growing." While the marble model would grow larger, there would be an increasing number of places for additional marbles to be placed, but at the beginning it would probably not be possible to place them at the rate of 60 marbles per minute. Thinking in terms of the marble model we have built and of the imaginary model, we might make the following predictions: During slow formation of crystals a few large ones will form. When the rate of growth is rapid, however, the excess solute cannot find its way to the original crystals, and many smaller crystals form. We now have predictions we can test.

Testing the Predictions With Sulfur Crystals

4 1

To grow sulfur crystals, you will need some small pieces of roll sulfur, 20 ml of carbon disulfide, a test tube, a stopper for the test tube, two watch glasses, and a 150-ml beaker. Put the carbon disulfide in the test tube and add some small pieces of the roll sulfur. Stopper the test tube and shake it to aid dissolving. Place about 5 ml of the solution of sulfur and carbon disulfide on one of the watch glasses and allow it to evaporate quickly. Place the remaining 15 ml of solution in the beaker and place

What effect will the rate of cooling have on the size of the crystals that form?

the second watch glass on the beaker so that evaporation is slower. Allow the beaker to stand overnight in a cool place. The next day examine the crystals formed in the beaker. [3] How do these crystals compare with those that formed on the watch glass? [4] Does this experiment confirm the predictions made on page 77?

Testing the Predictions With Copper Sulfate Crystals

4 2 Let's look at the growth rate of crystals in a copper sulfate solution. You will need 20 g of crushed copper sulfate crystals, a source of heat, a test tube holder, a stirring rod, one 20-by-150-mm test tube, three 13-by-100-mm test tubes, 25 ml of hot water, two beakers—one full of ice, a magnifier, and a test tube rack. Dissolve as much as possible of the crushed copper sulfate in 25 ml of hot water in a large test tube. Keep warming the mixture to aid dissolving, but *do not boil it.* Pour equal amounts of this solution into three test tubes. Let the first sample cool to room temperature by putting it aside for a while; cool the second sample very slowly by putting it in a beaker of warm water; cool the third sample by quickly nestling it in a beaker of ice. Study the size and shape of the crystals formed by the three different rates of cooling. [5] Were your results the same as those obtained by your classmates? [6] Were your results in keeping with the predictions we made on page 77? [A]

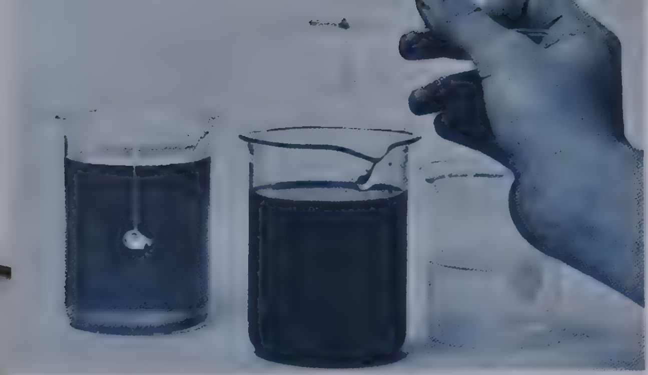
Pick out a single crystal for suspension in a test tube containing saturated alum solution.

4–3 OTHER TESTS OF THE PARTICLE THEORY

We know that supersaturated solutions are unstable. Because of their instability, the excess solute is apt to separate from the solution suddenly and sometimes unexpectedly. They can also be difficult to prepare. (You may have had difficulty when you tried to prepare a solution of hypo.) [7] Can separations be explained in terms of a particle theory? [8] Can you think of other ways to make the solid appear in a supersaturated solution? [9] Does anything in our theories explain the observed release of heat during this crystallization?

We could test our theory in another way, by trying to predict the most favorable conditions for the growth of large crystals of a particular kind of substance from solution.

4 3 You will need about 5 g of alum, 25 ml of water at room temperature, a stirring rod, one 20-by-150-mm test tube with a stopper, nylon thread, and a watch glass. Measure out 25 ml of water in a test tube. Dissolve as much of the alum as possible in the water, but do not heat the solution. When you are satisfied that the solution is saturated, pour about 5 ml into a watch glass. Stopper the test tube and save the remaining solution for later use. Set the dish of alum solution where it will not be disturbed. As the water evaporates, small crystals of alum will appear in the dish. Within 24 hours you should find one or two



Will chrome alum particles and copper sulfate particles grow on aluminum alum crystals?

crystals large enough to handle conveniently; these may be used as seed crystals. Remove them from the glass and gently dry them with tissue or a paper towel. Select one of the seed crystals and use the nylon thread to suspend it in the test tube of saturated alum solution. Do not stopper the test tube, but place it where the temperature is constant. Each day carefully remove and dry the crystal with a tissue. Then determine and record the mass, and return it to the test tube for further growth. Continue this activity for a week.

From our study of solutions, we believe that the solute is fairly evenly distributed throughout the solvent. As the solvent evaporates, the solution becomes saturated. As more solvent evaporates, crystals of solute appear. If the rate of evaporation is constant, then the rate at which the solid grows should also be constant. Thus, the amount of solute deposited on the crystal each day should be the same. [10] Do your results support this reasoning? The alum crystal is the first crystal we have prepared that closely resembles one of the models we constructed with marbles. While we cannot say that the particles in such a crystal are the same shape as the marbles, we can say a little about the arrangement of the particles in the crystal. If the marbles you used were not uniform in size, you may have noticed that you could not build regularly shaped models. Thus, we can conclude that all the particles in a crystal have the same size and shape.

Are there two different substances that have particles enough alike so that one may grow on the seed crystals of the other?

For a simple test to help you answer this question you will need a seed crystal of aluminum alum (potassium aluminum sulfate), nylon thread, about 6 g of chrome alum (potassium chromium sulfate), two 20-by-150-ml test tubes with stoppers, a test tube rack, a saturated solution of copper sulfate, and 25 ml of water. Prepare a saturated solution of chrome alum. Tie the seed crystal of aluminum alum to a nylon thread and suspend it in the saturated solution of chrome alum. Leave it there for several days. [11] Does the particle theory allow us to predict whether the particles of chrome alum will penetrate the clear alum crystal?

Which substance grew on the aluminum alum crystal?

As a second experiment put a seed crystal of aluminum alum in a saturated solution of copper sulfate and leave it there for several days. [12] From a comparison of alum crystals and copper sulfate crystals, can you predict whether copper sulfate will grow on the alum crystals?

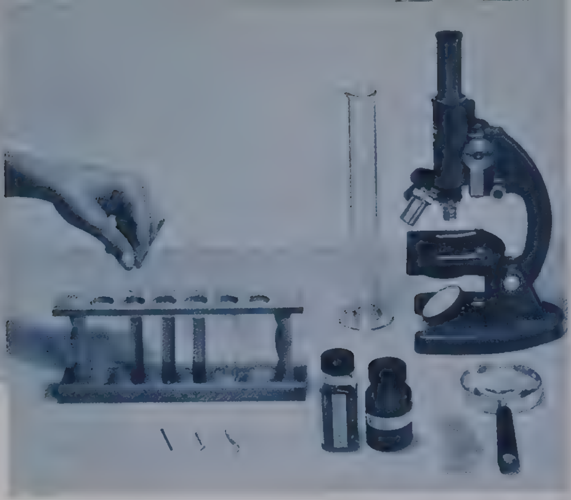
After several days use a knife blade to slice or break open the crystal suspended in the chrome alum. [13] What do your results indicate about the two kinds of alum crystals? Examine the alum suspended in copper sulfate solution. [14] What do your results indicate about the crystals of these two substances? [B, C]

4–4 METALLIC CRYSTALS AND THE PARTICLE THEORY

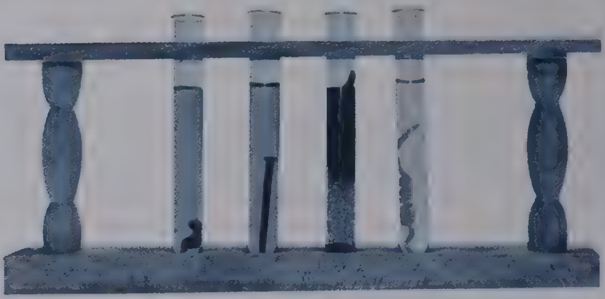
Under most circumstances metals seem to be continuous, but careful examinations under certain conditions reveal that metals are crystalline. A broken steel shaft may show a pattern that somewhat resembles the needlelike sulfur crystals we made in Chapter 2. By properly heating metals for an extended period, the small crystals which are normally present may grow to a relatively large size. Then etching away the surface with an acid reveals the crystalline forms. Because of the high temperatures required, this process for growing large metallic crystals is not possible in school laboratories. However, metallic crystals can be obtained from solutions of compounds that contain the metals.

To obtain metallic crystals you will need 5 g of crushed copper sulfate, silver nitrate crystals, 25 ml of water, four 13-by-100-mm test tubes, a 150-ml beaker, a test tube rack, a magnifier, a microscope, eight glass slides, an iron nail, a small piece of mossy zinc, 5 cm of magnesium ribbon, 10 cm of fine copper wire, and 10 ml of distilled water.

Prepare 25 ml of copper sulfate solution in the beaker by dissolving 5 g of the solid in water. Divide the solution equally among three



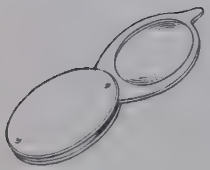
Growing metallic crystals from solutions of their compounds in contact with metal.



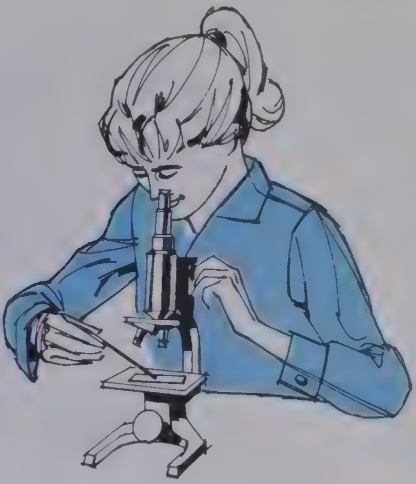
of the small test tubes. Label the test tubes *A*, *B*, and *C*. Drop the nail into test tube *A*, the magnesium ribbon into test tube *B*, and the mossy zinc into test tube *C*. Dissolve the silver nitrate crystals in the distilled water. Coil the copper wire by wrapping it around a pencil and then drop it into the silver nitrate solution. Without disturbing the contents, observe the metals closely and as often as you can over a period of 24 hours. Watch for color change in both the solutions and the metals. Record all observations in your laboratory record book.

When no further changes seem to occur, remove the metals from the test tubes and examine their surfaces. Put samples of each on the glass slides. You should observe these metals using different degrees of magnification: the naked eye, a magnifier, and a microscope at low magnification. In your laboratory record book, sketch the crystal pattern observed in each sample at each magnification. List their differences and their similarities. Now put a drop of each solution on separate glass slides. Let the water evaporate from these glass slides and then examine the residues. $\boxed{15}$ Do they contain evidence of crystals?

Preparation of metallic crystals by the methods used here involves *chemical reaction*, whereas preparation of crystals by the evaporation method involves only a physical change. In a chemical reaction, a *new substance* is formed which was not present when the experiment began. The iron in the nail, the magnesium, and the zinc are chemically more active than copper. Copper in turn is more active than silver. The more active metals replace the less active ones. The copper and silver appear



Teasing the drops of solution to cause precipitation.



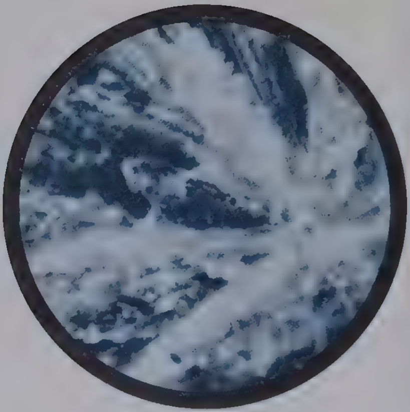
as free crystals, while the iron, magnesium, and zinc unite with the sulfate to form a new substance. We will say more about chemical changes when we study chemical energy in Chapter 11. The question is whether metals are continuous or composed of tiny particles. [16] What is your conclusion after examining these metallic substances?

4–5 CRYSTAL FORMATION BY PRECIPITATION AND FREEZING

The mixing of two different solutions will often result in the formation of a new substance which appears in the form of a crystalline solid. If this new solid does not seem to dissolve in water, it is called a *precipitate*.

It is most exciting to study the formation of a precipitate through a microscope.

You will need three clean glass slides, a clean stirring rod, a magnifier, saturated solutions of magnesium sulfate, barium chloride, strontium chloride, a concentrated solution of calcium chloride, and a microscope. Dip the stirring rod into the solution of calcium chloride and touch it to a glass slide so that a tiny drop of solution is deposited on the slide. Rinse the rod in water and dry it. Now dip the clean rod into the saturated solution of magnesium sulfate and touch it to the same slide so that the two drops of solution are very close together. Put the slide on the stage of the

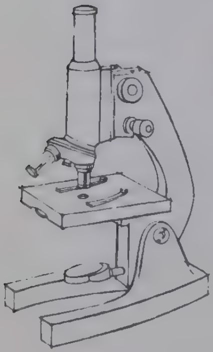


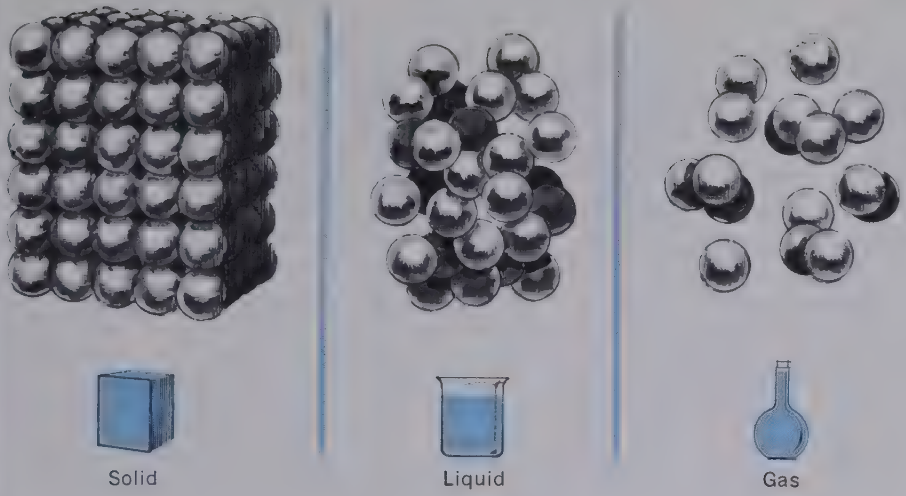
Microphoto of benzoic acid crystals. These crystals form in a manner similar to salol.

microscope and focus on it. While you look through the microscope, tease the two drops of solution together with your stirring rod. Study the resulting crystals of calcium sulfate carefully and sketch their pattern in your laboratory record book. Continue to observe these crystals until most of the water has evaporated. [17] How does the appearance of the crystals change as the water evaporates? [18] Which theory of matter would you use to explain these observations?

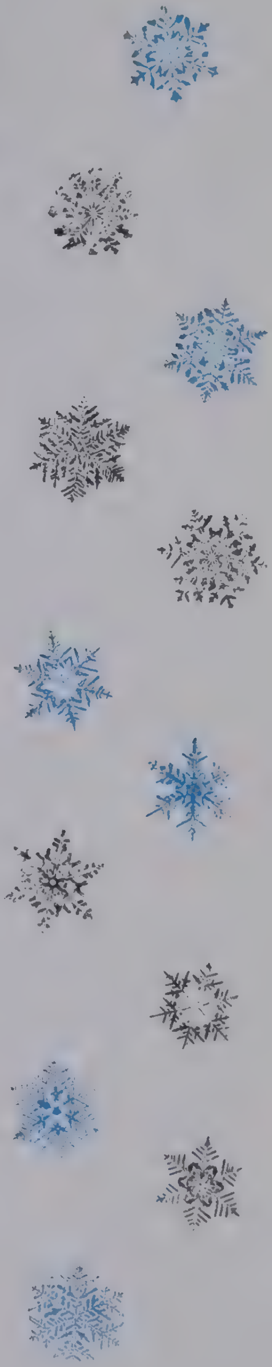
Repeat the process on another glass slide, using a drop of saturated barium chloride solution and a drop of magnesium sulfate solution. The new solid formed here is barium sulfate. Sketch this crystal in your laboratory record book and label it. Then repeat the process once again using a drop of saturated strontium chloride solution with the magnesium sulfate solution. The crystals formed are strontium sulfate. Sketch these in your laboratory record book and compare the shapes of the crystals you prepared by precipitation. (Be sure to label each entry in your notes; a good scientist does not like to trust his memory.)

4 7 You have seen ice crystals melt into water many times and you know that water can be refrozen, or recrystallized. You have also seen sulfur melt and recrystallize. Salol (phenyl salicylate) melts and recrystallizes at a more convenient temperature than ice or sulfur. Obtain some salol and put a few crystals on a glass slide; heat the slide gently to melt the crystals and then let it cool. As the slide cools, use a microscope or a magnifier to observe the formation of crystals. [D]





Illustrations depicting the three states of matter.



4–6 CRYSTAL PATTERNS

Now let's study crystal patterns more carefully. Recall the frost crystals that formed on the container of salt and ice in Chapter 2. [19] Were their patterns similar to any of the other crystal patterns you have studied? [20] By what method were these crystals prepared? Some of nature's most beautiful patterns are found in frost and snow crystals. You may want to do some reading on the atmospheric conditions that influence their formation.

Minerals furnish interesting examples of crystal patterns. Examine samples of several kinds of rock and minerals to determine how their crystal patterns are structured. [21] How can the variety of patterns be explained? [22] Is the particle theory a source of explanation?

Carbon is a familiar material. Is it crystalline? Graphite is a form of carbon used in making pencil "lead." Examine some graphite with a microscope and describe its pattern. Also scrape a bit of carbon from a burned wooden match. Examine the carbon under a microscope and compare it with graphite. [23] Does it look the same?

Remember that the degree of magnification affects our observations. Therefore, we must be very careful in stating conclusions about the crystalline or noncrystalline nature of a sample of matter. Examine many different common substances. [24] Do you find that crystals are a common form for matter? [E, F, G]

4–7 CLASSIFICATION OF CRYSTALS

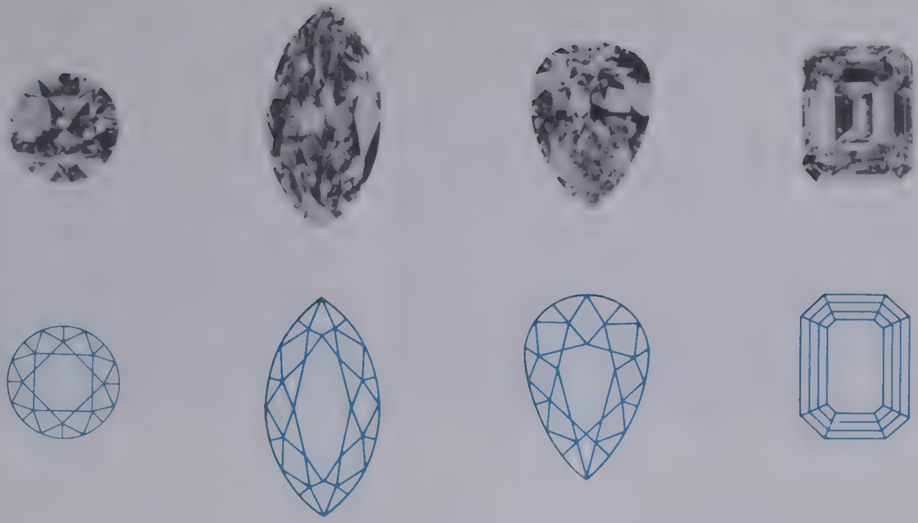
It is fun to grow and study crystals, but our real purpose in preparing many different kinds of crystals by using a variety of methods has been to gain a better understanding of the nature of matter. We have also been attempting to answer questions by following the procedures used by scientists. You will recall that one of these procedures is classification. Development of a good system of classification often enables us to see new relationships.

From your laboratory activities you have acquired many kinds of crystals and obtained information about still more of them. It would be most helpful to you to devise a system of classification to use in grouping your samples, including the crystals you have sketched in your laboratory record book as well as those you have grown. Color, shape, method of preparation, and size will need to be considered as you develop your system of classification. [H, I]

4–8 CRYSTAL PATTERNS AND THE NATURE OF MATTER

You have watched ice melt and become liquid; you have observed solids in the process of dissolving. You have noted that both of these processes occur more rapidly if heat is added to the samples. You have learned how to arrange conditions so that a dissolved substance deposits on a seed crystal to produce a larger crystal. You have used spheres and cubes to build models that seem to resemble crystals. Throughout all these many activities one fact has become increasingly clear—matter does seem to be made of particles. Whether a given substance is in the solid state seems to depend upon the orderly placement of tiny pieces. Recall the dissolving of a sugar cube. A regular, six-sided lump of sugar slowly but steadily disappeared. The single crystals moved apart until they were finally completely dispersed in the water. When the closely packed arrangement of the cube had disappeared, you could no longer identify the bits of sugar. Even with a high power microscope the sugar was invisible.

As you heated the crystal model of marbles and wax, you noticed that when the wax melted the well-arranged marbles fell apart. When the "dissolving" process was complete, the marbles were scattered about on the bottom of the beaker. The main characteristic of crystalline solids is orderliness that can be reproduced after the solids have been subjected to melting or dissolving.

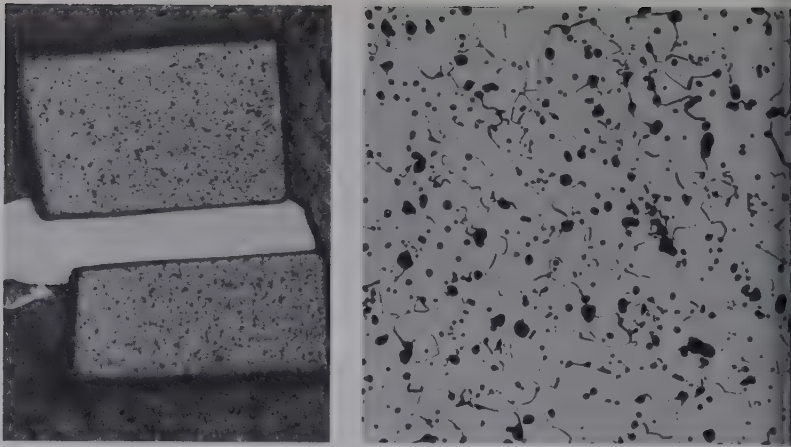


The four basic cuts of diamonds, a crystal form of carbon. Cutting gem stones such as these requires detailed knowledge of crystal structure.

Another observation that we have made is that each substance has its own distinctive crystal pattern. You saw evidence of this when you placed an alum crystal in solutions of chrome alum and copper sulfate. The chrome alum grew on the aluminum alum and the copper sulfate did not. The particle theory allows us to formulate an explanation: The particles of chrome alum can fit into the pattern of the aluminum alum and the particles of copper sulfate cannot. Examination of alum crystals and copper sulfate crystals shows at a visible level that the two substances have different crystal patterns. [25] In your classification of crystals, did you place them in different groups?

Then the evidence from our study of crystals seems to tell us that matter is made of particles and that each different kind of matter has a different pattern for the arrangement of these particles. When the particles are very close together and assume an orderly arrangement, a solid such as ice exists. When the particles are moved apart, the pattern is lost and a liquid exists. If these particles are moved still farther from each other, a gas exists. We will investigate these ideas again in our study of energy.

In general, matter has a solid, a liquid, and a gas phase. Which of these forms the particles take seems to depend on the energy they possess. [J, K]



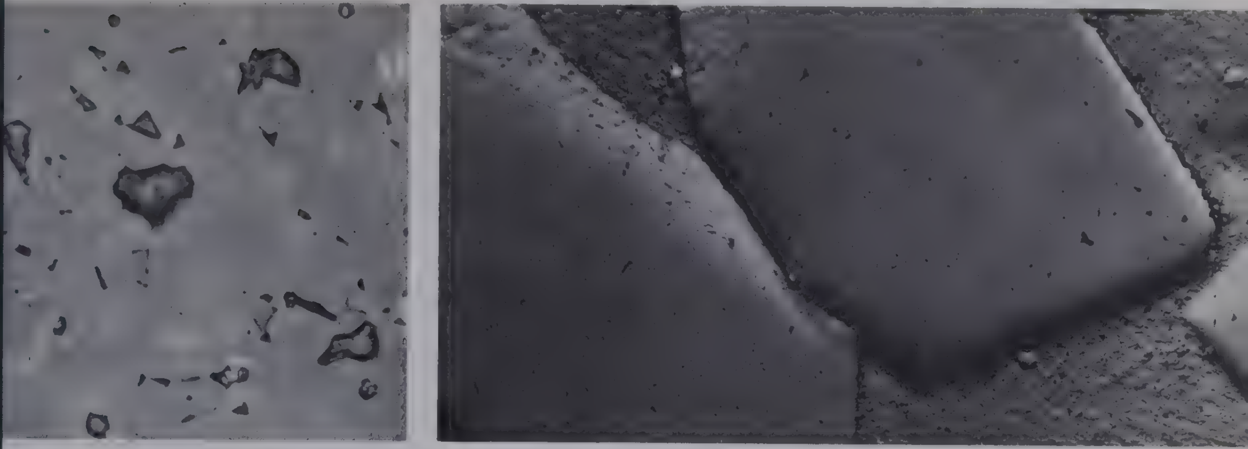
From left to right – aluminum oxide magnified 18 times; magnified 180 times; magnified 900 times. The fourth photo shows a carbon replica of aluminum oxide magnified 12,000 times.

4–9 VARIATIONS IN CRYSTAL PATTERNS

Now that you have had experience with solutions, with crystal formation under various conditions, and with possible models of crystals, let's return to a question raised in Chapter 2. [26] What causes sulfur to form into rhombic crystals as the solution of carbon disulfide evaporates, into monoclinic (needlelike) crystals when molten sulfur is cooled in a filter paper, and to exhibit no visible crystals when hot sulfur is poured into a beaker of cold water? You may be able to answer this question now. Carefully examine the crystals in the cone of sulfur that you made in Experiment 2–2. Compare the color and the structure of these crystals now with the description you wrote in your record book when they were formed. Also examine the sample of amorphous sulfur from Experiment 2–2 and note how it has been altered. [27] Will the amorphous sulfur still float in tetrabromoethane solution? [28] What factors might be responsible for the changes that have occurred in these samples of sulfur?

[29] Why did larger crystals of copper sulfate form in the sample that was allowed to cool very slowly, while those which formed rapidly by plunging the solution into ice were small?

Answers to these questions can be formulated if you will think about the crystal models you built. [30] What was required to make a good model from marbles and cement? It took much time and patience on your part to assemble the pieces of material (marbles) into an orderly pattern. Could you have achieved the same result if you had quickly tossed the marbles and cement together? Obviously not. The tiny particles of solute in a solution must be given adequate time if they are to become arranged into a characteristic repeated pattern. Rapid cooling of a molten sub-

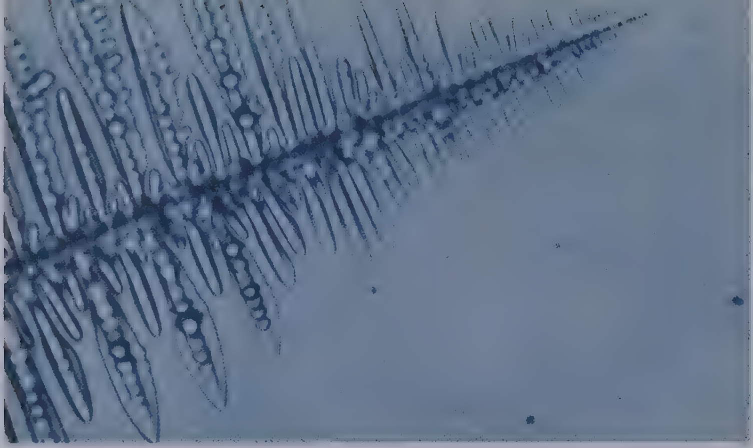


stance or of a saturated solution will result in hasty assembling of the particles. Slow and gradual cooling allows the small pieces of matter sufficient time to be arranged into the more perfect patterns we usually associate with crystals.

4–10 VARIETY OF PARTICLES

As you look around you, it becomes readily apparent that matter has tremendous variety. Thousands of different substances exist. How does this variety fit into a particle theory? The answer is that each kind of substance is made of its own unique particles. The particles are of two kinds, simple and compound. The simple particles are *elements*. There are about 100 elements that have been identified by scientists. You are familiar with many of them, such as iron, carbon, aluminum, and silver. The elements may be classified as *metals* or *nonmetals*. Metals are usually characterized by their luster, hardness, and ability to conduct heat. Nonmetals do not have these properties. Many metals have names ending with the suffix *-ium* (sodium, calcium, magnesium, chromium). However, there are some well-known exceptions: copper, gold, silver, and mercury.

When the particles of elements unite to form more complex substances, the new substances are known as *compounds*. Usually the name of a compound is determined by the names of the elements from which it is formed. The name of the metal is given first and the name of the nonmetal last. If only *two* elements are present in the compound, the name of the nonmetal is changed so that it ends in the suffix *-ide*. For



Microphoto of cyclohexanal crystals, which can be grown in a manner similar to salol.

instance, you have used carbon disulfide for dissolving sulfur. The suffix *-ide* tells us that only two elements, carbon and sulfur, are present in the compound. The prefix *di-* tells us there are two parts of sulfur for every one part of carbon.

The suffix *-ate* (as in copper sulfate) indicates the presence of more than two elements in a compound. Usually one of the elements is oxygen. Thus, the name "copper sulfate" tells us that the elements copper, sulfur, and oxygen are present. Copper, nitrogen, and oxygen are present in copper nitrate.

Rules for naming other compounds are more complex than those we have mentioned, but with these guidelines it should be easier for you to understand the names you will encounter in your present studies. [L]

TAKING INVENTORY

Through your crystal-growing activities you have learned to test a scientific theory. Since the particle theory has provided a basis for explaining many different natural occurrences, it has become increasingly acceptable.

In addition, you have learned through a number of activities that crystals may be formed by

1. Evaporation of solvent from a solution
2. Disturbing the equilibrium in a supersaturated solution
3. Precipitation during a chemical change
4. Cooling of a melted substance

The rate of crystal formation has a great influence on the size and perfection of the final specimen. This is true both in the laboratory and in nature.

The orderly arrangement of crystals and the disorder that occurs as crystals melt or dissolve lend much support to the theory that matter is made of particles.

A study of minerals reveals that crystals formed under natural conditions are much the same as those we have made. Classifications of both crystals and minerals depends on a rather thorough knowledge of these substances. As in button grouping, the purpose of the classifying determines the system to be used.

The evidence we have acquired from our investigations seems to favor a model of matter consisting of very similar particles which are extremely small.

Throughout our study of solutions and crystals, heat has been an important factor. Up to this time we have not considered *why* heat has been important, nor have we asked *what* heat is or *how* it is able to make such a difference in the ability of matter to change from a solid to a liquid to a gas. We are now ready to try to answer these questions. The next several chapters will take us into investigations that will reveal much about the kinds of energy—the cross threads which make up the fabric of our universe.

FURTHER EXPLORATIONS

A. Prepare small crystals of copper sulfate, Rochelle salt, nickel sulfate, and potassium ferricyanide from water solution to obtain a variety of patterns and colors to study. You can also use these crystals as seed crystals for growing large samples.

B. Can you insert crystal seeds of two solutes in a saturated solution containing both solutes and obtain crystals of both?

C. In our study of solutions we have assumed that the solute is evenly distributed throughout the solvent. Is this an acceptable assumption? What happens to the uniformity of a solution if it stands undisturbed for a long time? What happens to uniformity as solvent evaporates from the surface of a solution? Your study of solutions and crystals should give you ideas on how to investigate the uniformity (or lack of it) in solutions. Choose one aspect of the general question and investigate it as a project.

D. The following method is another way to make crystals by precipitation. Put distilled water in a petri dish to a depth of 6 or 7 mm. Then drop crystals of silver nitrate and sodium chloride into separate places in the dish. Observe the reaction. An overhead projector can be used for this purpose. Is this experiment another source of information about the nature of matter?

E. Investigate the nature of stalagmites and stalactites which are found in caves. Are these formations crystalline? How are they formed?

F. Diamond is a form of pure carbon, although it is quite different in appearance from graphite. Consult a reference book to learn what is responsible for this great difference.

G. In your study of crystals found in nature, you have observed that their sizes vary considerably. Propose a hypothesis to explain these differences and consult a book on minerals to check your guess.

H. If you are especially interested in classification of crystals, you may study the accepted systems of classification and use these in grouping the crystals you have grown.

I. Select several crystals of any given substance and determine if the angle between any chosen pair of faces is the same on all the crystals. In order to measure such an angle you will need to construct a *goniometer*. Directions for making this instrument are given in the book *Crystals and Crystal Growing* (see Further Reading below). After you have learned to use the goniometer, you may want to measure the interfacial angles on a variety of crystals. Record the data in your record book and decide if this is a reliable method for identifying crystals.

J. Can you distinguish grains of salt from grains of sugar by examining them with a magnifier or a microscope? How could you separate salt from sugar if they were accidentally mixed, without losing any significant amount of either one?

K. You may become sufficiently interested in crystals and their forms to make a hobby of studying them. Simple and inexpensive models may be obtained from Mr. A. J. Gude, 845 Dudley Street, Lakewood, Colorado. Write to him for information.

L. Identify the chemical names for the various compounds which, up to this point, you have had occasion to use. Use the names to judge the number of elements each contains.

FURTHER READING

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A technical but clearly written book on crystal growing. Contains many illustrations of apparatus and photographs of crystal surfaces.

Bunn, Charles W., *Crystals: Their Role in Nature.* Academic Press, Inc., New York, 1964.
This book attempts, in an informal way, to explain the scientific field of crystallography.

Holden, Alan, and Phyllis Singer, *Crystals and Crystal Growing.* Anchor Books, Doubleday & Company, Inc., Garden City, N.Y., 1960.
An excellent source of information on how to grow crystals.

Lonsdale, Kathleen, "Disorder in Solids." *Chemistry*, Vol. 38, No. 12 (December, 1965), pp. 14–19.
An interesting discussion of the solid state. The article also discusses x-ray diffraction techniques for obtaining information about crystal structure.

PEARL, RICHARD M., *Rocks and Minerals.* Barnes & Noble, Inc., New York, 1956.

Tells how to recognize and identify minerals, rocks, ores, metals, and crystals. Discusses how these specimens are classified, how to collect them, and how to display them.

———, *How to Know the Minerals and Rocks.* New American Library, New York, 1957.

Describes identification keys for minerals and rocks. Tells where various specimens can be found.

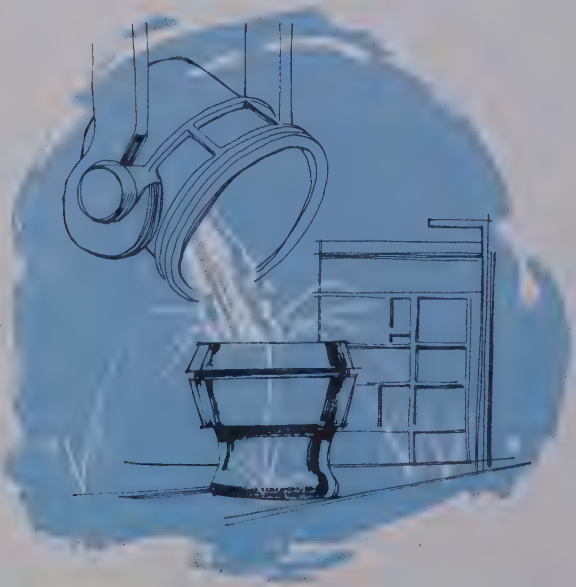
WOHLRABE, RAYMOND A., *Crystals.* J. B. Lippincott Company, Philadelphia, 1962.

Contains many suggestions for exploring the world of crystals.

ZIM, HERBERT S., and PAUL R. SCHAFFER, *Rocks and Minerals.* Golden Press, New York, 1957.

A guide to identifying rocks and minerals; includes color illustrations. Bead tests, flame tests, the hardness scale, and methods for determining specific gravity provide interesting activities for students interested in rocks and minerals.

5 Temperature and Heat



5-1 Intensity and Quantity—Two Characteristics of Heat
5-2 The Need for Thermometers
5-3 Measuring Temperature
5-4 What Is Heat?
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5-7 Measuring Heat
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5-9 Heat-Holding Ability of Solids

The nail does not hold as much heat as the water. How many times would you need to heat the nail and dip it into the water to cause an appreciable change in the temperature of the water?

5-1 INTENSITY AND QUANTITY—TWO CHARACTERISTICS OF HEAT

We have found that temperature is a factor in the changes we observed in many of our experiments. With salts, a greater quantity dissolved and the dissolving process was faster in hot water than in cold. Heat causes ice to melt and water to boil. Certain solids, such as ammonium chloride and hypo, lowered the temperature of the solution as they dissolved. Now let's explore temperature and heat and the part they play in so many changes in matter.

Direct evidence may help us understand the relationship between heat and temperature. You will need a can or beaker of about 250-ml capacity, 200 ml of water, a source of heat, a short strip of metal (such as copper wire or a nail), a pair of pliers or forceps, an asbestos pad, a ring stand and ring, and a piece of paper. Heat the 200 ml of water in the can or beaker over a moderate flame. While the water is heating, grasp the metal strip with the pliers and hold it in the flame. When the metal acquires a red glow, remove it from the flame and hold it against the piece of paper. That the metal burns a hole in the paper shows that it is very hot. The flame has been warming the container of water for a longer time than was needed to heat the metal, yet we know that neither the container nor the water will burn the paper. Which is hotter, the water or the metal? It would seem that we cannot measure heat by simply determining how hot an object becomes when it receives that heat.

Temperature is the degree of hotness, but heat seems to be something different. Temperature measures the *intensity* of heat—the more heat in

Are the temperature sensations we feel good indicators of the actual temperature?

any one object, the higher the temperature; the less heat in any one object, the lower the temperature. In the metal strip, heat was concentrated in a relatively small object. Because of this concentration, the temperature of the metal was much higher than the temperature in the container of water. A larger quantity of heat was added to the water, but because it was distributed into a greater amount of material, the temperature was not so intense.

5–2 THE NEED FOR THERMOMETERS

We cannot rely on our senses to report the condition called temperature with any accuracy. We are influenced by temperatures to which we have been exposed.

5 2 For a test of your ability to sense temperature you will need three sizable containers (at least 1000 ml beakers) and quantitites of hot, cold, and room-temperature water. Put equal amounts of the hot water, room-temperature water, and cold water in the three containers. Insert one hand in the hot water and the other hand in the cold water. Leave your hands in these containers for 30 seconds. Now put both hands in the third container. $\boxed{1}$ What do you notice? Certain parts of the body, such as the inner wrist, are useful as temperature indicators, but we need instruments for a more accurate measurement of temperatures.

Table 5–1

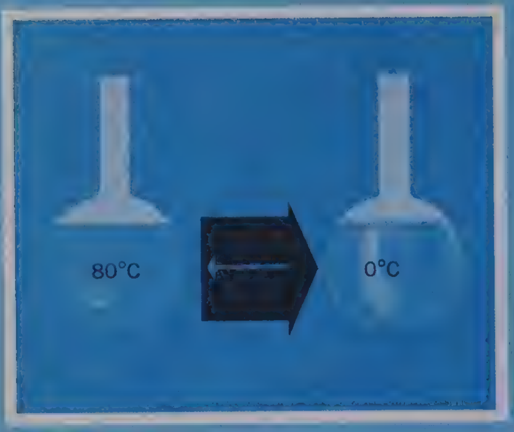
Cricket chirp rate on a summer night in Reno, Nevada.

Temp. °C	Chirps per sec.
26	1.93
26	1.93
26	1.95
26	1.92
26	1.99
25.5	1.95
25.5	1.93
25.5	1.95
25	1.86
23	1.76
20	1.50
20	1.50
19	1.40
18	1.36
13	0.57
13	0.65
13	0.94
13	0.84
13	0.84
10	0.58

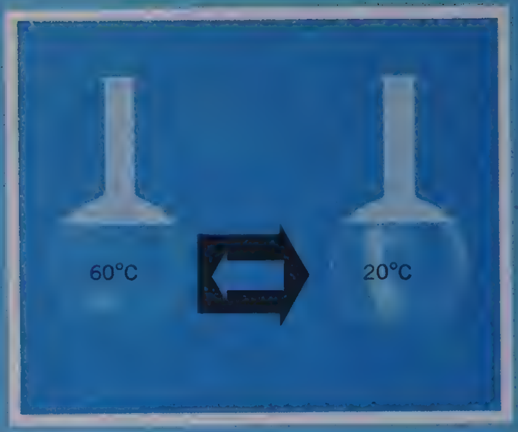
Man is said to be "warm-blooded" because he maintains a relatively constant body temperature regardless of the temperature of the environment. All mammals and birds are warm-blooded. All other animals, including insects, are said to be "cold-blooded" because their temperature changes with that of their environment. As body temperature changes, the rate of activity changes. You may have noticed that the chirping rate of crickets is related to the temperature. As the temperature drops, their chirp rate and body motions become noticeably slower.

In late summer, one of the authors took some data on the chirp rate of tree crickets in some honeysuckle, using a tape recorder to be sure the results could be reexamined and measured accurately. Table 5–1 shows the average rate in chirps per second and the temperature as the sounds were counted and timed. Prepare a graph on which the rate (number of chirps per second) is plotted as a function of temperature. Does your analysis of the data indicate a consistent relationship between the chirp rate and temperature? On an evening when several crickets are serenading, you have probably noticed a slight difference in rate among members of the chorus. [2] Would you guess that this is the result of individual differences among the crickets or the slight differences in air temperature around the crickets? [3] Through what range of temperature do you think these insects might be reliable "thermometers"? [4] Would you expect the chirp rate to depend on the species of cricket? on the geographical location? on the age or maturity of the insect? [A]

Imagine that the flasks are suspended and enclosed in an insulated box.



Heat exchange is very unbalanced between two flasks if there is a great difference in the temperature of their contents.



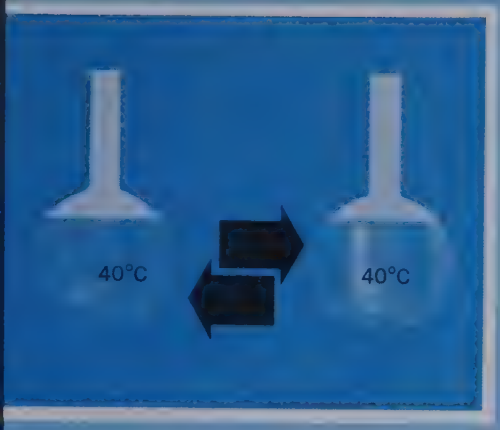
The heat exchange becomes more balanced as the temperature differences decrease.

5–3 MEASURING TEMPERATURE

There is an important principle that must be established before we can measure temperature with any device: *Every object will come to the same temperature as its surroundings if it is left long enough.* A thermometer only reads the temperature of a particular region reliably when it comes to the same temperature as that region.

Perhaps this principle is related to the reason we cannot readily or accurately determine temperature with our senses. Our nervous system responds to exchanges of heat rather than to the actual temperature. Experiment 5–2 with hot and cold water should help you realize this. Your impression of temperature was related to the fact that the water temperature was different from the temperatures to which your hands had become accustomed.

[5] If our senses are not dependable for temperature determination, what may we use for measurement? When we see ice forming in puddles we know that the air temperature is at the freezing point of water. When the ice melts we know that the temperature is going above that point. Therefore, water can be used as a crude indicator of temperature, but only within a narrow range. We must find some property of a substance that will allow us to measure a wide range of temperatures on a scale.



Heat exchange is balanced when the temperatures are the same. The system is at equilibrium.

Scientists have already given us several possible solutions to the problem. We do not know how long ago it was that man discovered that the volume of a material increases with increasing temperature. We do know that Galileo, working nearly 400 years ago, experimented with air thermometers. Newton, in the seventeenth century, made a liquid-expansion thermometer similar in principle to our present instruments. We may use this relation of volume and temperature of a liquid to establish a number scale for regions of temperature. If liquid mercury is sealed in a long narrow tube, the position of the end of the liquid column indicates the volume and hence the temperature of the liquid in the bulb.

You can build a simple thermometer like the one shown in the photograph above and with it you can study the relationship between the temperature and the volume of a liquid.

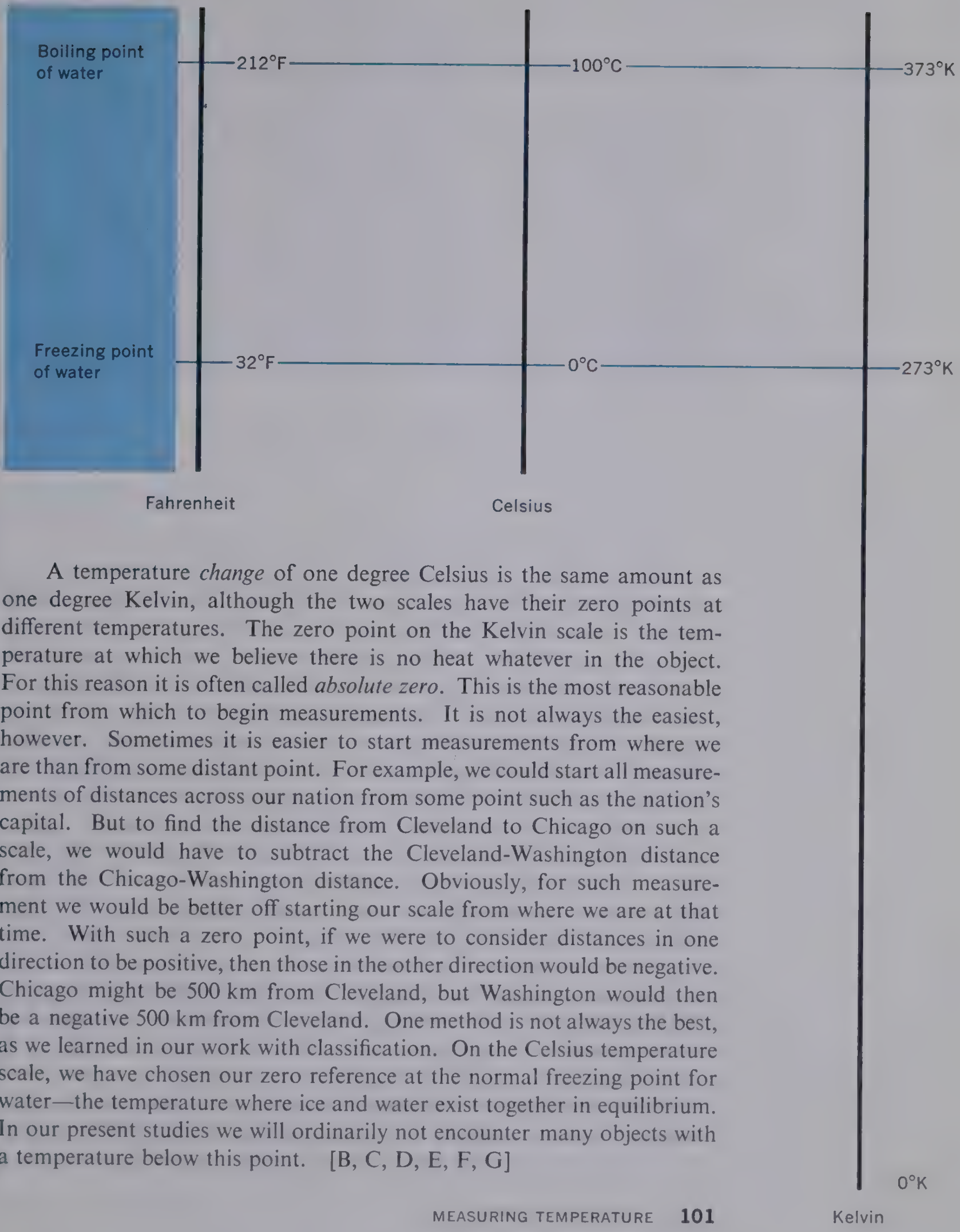
You will need a test tube fitted with a one-hole stopper and filled with water, a 30-cm length of glass capillary tubing, a candle, a 250-ml beaker, a ring stand and ring, an asbestos pad, and a metric ruler. Moisten the tubing and stopper. Carefully insert the tubing into the stopper hole. (**CAUTION:** Be sure to follow your teacher's instructions in carrying out this operation.) Fit the stopper and tubing into the full test tube and put the test tube in a water bath above the flame; do not allow the water in the test tube to boil. With the metric ruler, measure the length of the column of water above the stopper every 30 seconds until the water has expanded to fill most of the tubing. Make a graph plotting the length of liquid column (vertically) as a function of time of heating (horizontally). Since we can assume that the inner diameter of the tubing is probably quite uniform, equal lengths along the tube represent equal increases in the volume of the liquid. (This may seem obvious to everyone, but it is important that we stop occasionally and describe the obvious to be sure we are not building our argument on false assumptions.)

There is another important assumption we should make at this time, namely, that in all probability the candle burned with a nearly constant flame. If heat is released at a rate that depends on the size of the flame, a large flame uses up the candle wax at a greater rate than a smaller flame and releases heat at a greater rate. If there is no flame at all, no heat is being released. Our candle flame, which was more or less uniform, released heat to its surroundings, including the water bath and test tube, at a nearly uniform rate.

The graph you just made of column length versus time now tells us much more. Since volume is proportional to length, and since heat is proportional to time, the graph can be read as the relationship between volume increase of the liquid and the heat absorbed. The device used in the experiment is indeed a thermometer; it operates on exactly the same principle as the thermometer you used in your study of solutions.

To make good comparisons of temperature we need to be able to express the temperature numerically. Numbers are much less ambiguous than words. As we have learned, our senses are not reliable for judging temperature and are of little value in experimental work. An accurate knowledge of the actual temperature is important in many activities, including cooking. Every homemaker knows that if a cake is to turn out well, the oven temperature must be kept constant.

For our temperature measurements in science we most often use two scales. One of these, named after the Swedish scientist Anders Celsius, is called the Celsius scale, and is recognized by the letter "C" after the temperature reading. Because it divides the temperature between the freezing and boiling points of water on the scale into 100 divisions, it is also called the centigrade scale. The name Celsius is preferred because centigrade, which means 100 graduations, could also be applied to the meter stick if we chose, as the centimeters are centigrade divisions of a meter. The other temperature scale used in science is called the Kelvin scale, named for Lord Kelvin, a British physicist; it is indicated by the letter "K." [6] Why do you think the Fahrenheit scale is not used in scientific work?



A temperature *change* of one degree Celsius is the same amount as one degree Kelvin, although the two scales have their zero points at different temperatures. The zero point on the Kelvin scale is the temperature at which we believe there is no heat whatever in the object. For this reason it is often called *absolute zero*. This is the most reasonable point from which to begin measurements. It is not always the easiest, however. Sometimes it is easier to start measurements from where we are than from some distant point. For example, we could start all measurements of distances across our nation from some point such as the nation's capital. But to find the distance from Cleveland to Chicago on such a scale, we would have to subtract the Cleveland-Washington distance from the Chicago-Washington distance. Obviously, for such measurement we would be better off starting our scale from where we are at that time. With such a zero point, if we were to consider distances in one direction to be positive, then those in the other direction would be negative. Chicago might be 500 km from Cleveland, but Washington would then be a negative 500 km from Cleveland. One method is not always the best, as we learned in our work with classification. On the Celsius temperature scale, we have chosen our zero reference at the normal freezing point for water—the temperature where ice and water exist together in equilibrium. In our present studies we will ordinarily not encounter many objects with a temperature below this point. [B, C, D, E, F, G]

5–4 WHAT IS HEAT?

We can now make two general statements about heat:

1. Temperature is a measure of the amount of heat in a given object.
2. Heat such as that from a burning candle causes some liquids to expand.

We observed that heat caused the expansion of water in our thermometers in Experiment 5–3, and we know that at least one other liquid—mercury—expands by heating. $\boxed{7}$ Why does heat produce such a result?

In our study of matter we found it useful to construct several imaginary models or theories about its structure. Then we compared each new observation with these models and selected the best possible explanations. Let's do the same for heat. A summary of our observations so far includes the following:

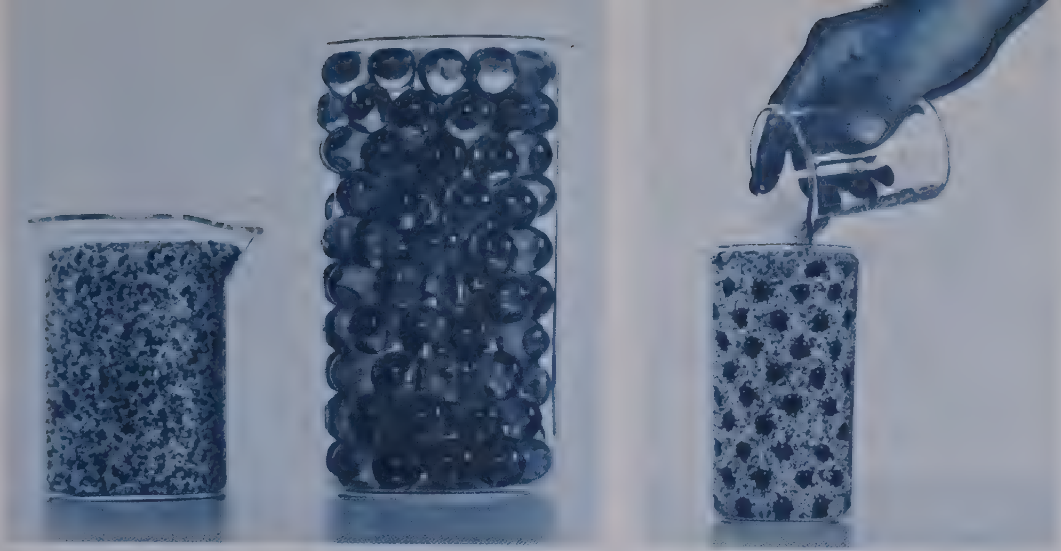
1. Adding heat causes objects to become hotter.
2. Heat causes liquids to expand.
3. Heat may change the state of matter from solid to liquid to gas.
4. Heat may be exchanged from one object to another. It seems to flow, and always from the warmer object to the colder.
5. Heat may be released or absorbed during certain changes in matter.
6. Heat acts upon our bodies, particularly certain parts of the nervous system.

The simplest way we can describe heat at this time is to say that it is something which flows from hot objects to colder objects, and from hot spots to cold spots within an object. In an object, heat produces changes. These may be changes in temperature and volume plus changes of state, such as solid to liquid or liquid to solid. Heat may also produce changes in living tissue, including signals in the nervous system.

Let's explore our surroundings while thinking of heat as a form of matter. We can try to learn more about the special properties of heat in order to explain its behavior. For convenience, we may refer to heat as a caloric fluid. The word *caloric* means heat.

$\boxed{8}$ Then is caloric fluid, which seems to flow like other fluids we have observed, really a form of matter? Our evidence indicates that matter is made up of particles, has mass, and occupies space. $\boxed{9}$ What properties would caloric fluid have to have to account for the changes we have produced?

Caloric fluid seems to have volume. As it flows into an object, it seems to take up space, much the same as when some other set of particles flow into an object. (Remember that when we added alcohol to



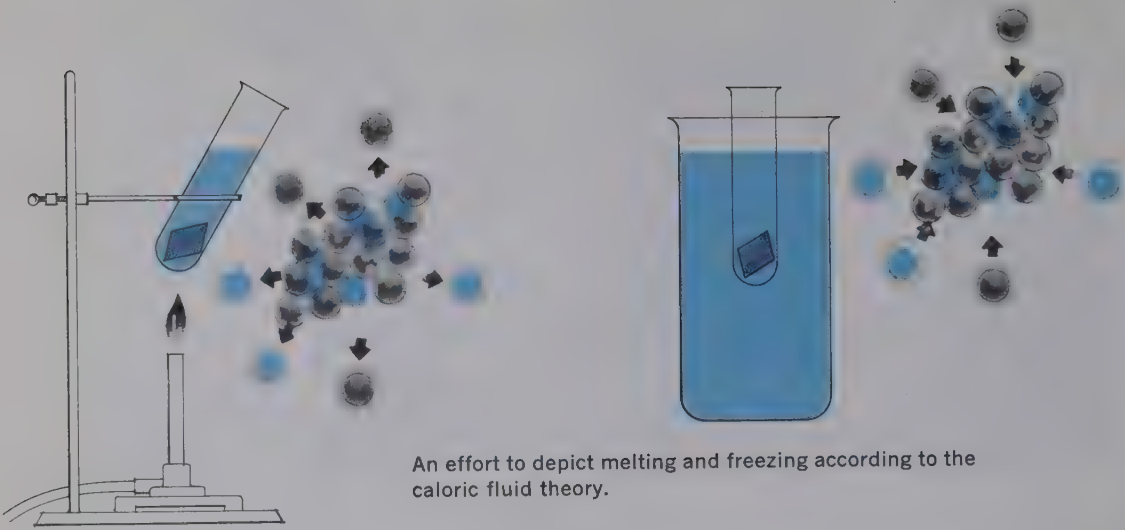
Is heat a substance? Does it flow into the spaces between particles and thus push them apart?

water it seemed as if the particles of one filled spaces between particles of the other). Adding a certain amount of heat to an object should therefore cause it to expand by that same amount. Experiment 5–3 with the simple thermometer seems to confirm this.

How can we explain this expansion? We may assume that the caloric fluid occupied spaces between the water particles and separated them slightly, causing the total volume to increase. We would get a similar effect if we packed all the crevices of a stack of marbles with sand, using a little more than they could hold. The stack would expand somewhat.

A model for heat requires that caloric fluid flow easily into the spaces between the particles of matter, causing these particles, which were originally attracted to one another, to move apart. Perhaps a property of caloric fluid is to reduce the attraction between particles of matter, or perhaps it is attracted to the particles of matter and causes the increase in volume by repelling itself. At any rate, from our marble model we can predict that solids will expand in the same general manner as liquids. (The expansion of solids is harder to measure, however, since we cannot get a large specimen to expand within a tube as we did with a liquid.)

Suppose we find that the amount of expansion differs from one substance to another when a given amount of heat is added. This difference may be because of different-sized spaces between particles. Another explanation might be that the attraction between the particles



An effort to depict melting and freezing according to the caloric fluid theory.

of the substances is different, and therefore, it is affected differently by the addition of the fluid. Also, solids do not change in size as much with each change of temperature as liquids do. Our thermometer would be useless if the glass bulb became larger at the same rate as the liquid it contained.

5–5 CHANGE OF STATE IN MATTER

If we apply the caloric fluid theory to the melting of sulfur and ice you observed, we may assume that the fluid runs through solid matter, penetrating the spaces between the particles. Gradually, the fluid surrounds each particle, reducing its attraction to neighboring particles. As this happens the solid begins to break into a less orderly arrangement, and finally the particles flow freely, assuming the familiar characteristics of a liquid.

Our study of crystalline solids has shown us that the particles of which they are made may be arranged in different patterns. This could mean that the particles making up different kinds of crystals have different shapes. If we check the melting temperature of various solids, we may find that the amount of caloric fluid needed to overcome the force of attraction between the particles is different for the different crystal patterns.

We can also explain, in terms of our theory, the process of changing from a liquid to a vapor, commonly called evaporation. When enough

of the heat fluid has been added to a liquid, the particles of the liquid will be surrounded: Their attraction for one another will be overcome and the particles will break completely from neighboring ones and move off independently. Thus the gas (vapor) would be expected to spread more or less evenly through the air.

5 4 We can investigate the effects of adding heat at a uniform rate to a substance such as water. You will need a small can (such as a tuna-fish can), a tripod support or a ring stand and ring, a source of heat, a thermometer, a stirring rod, several ice cubes, and paper towels or cloths. Blot the extra water from a few cubes of ice and put them in the can. Support the can just above the tip of the flame, which should be as uniform as possible. Stir the ice cubes so that heating will be uniform, and record the temperature at frequent intervals. Place the thermometer in the can so that its bulb does not rest on the bottom. If possible, continue to heat the can until the water has evaporated. We are interested in checking the amount of time needed to melt the ice and to heat and finally evaporate the water, as well as the rate at which the temperature changes. Plot a graph of temperature as a function of time. On your graph, mark the temperatures at which the contents are all ice, ice and water together, all water, and boiling.

A theory should not just explain what we have observed; it should also prove useful in predicting what we have not yet observed. More than casual observation is needed to establish that a theory is correct. To increase our confidence in a theory, we must measure and calculate to be sure our observations are consistent, regardless of the size of the sample. Try to predict the shape of the graph if twice as many ice cubes had been used. Perform the experiment, record and plot the results.

We believe that our physical surroundings do not "remember" as we do. For example, if what had recently happened to a sample of water affected its boiling temperature, we would have an interesting set of circumstances. It would be meaningless to suggest that something be boiled for a certain time if we did not know that, in general, there is a single boiling temperature. [10] Does water everywhere in the world boil at the same temperature at all times? [11] What are the conditions that determine boiling temperature? We might also wonder if all ice water is at the same temperature no matter where in the world it is measured. Have you ever measured the freezing—or melting—temperature?

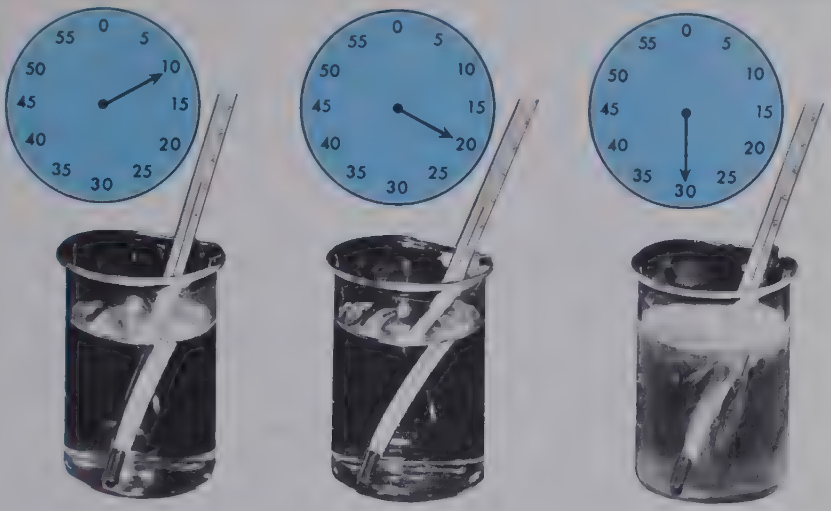
Although freezing temperature is not the same as *time* to freeze, we should expect that the recent history of a substance should not influence the time of freezing. If this is true, is there any truth to the saying that warm water will freeze faster than cold water? About 700 years ago Roger Bacon, one of the pioneers in experimental science, wrote,

> Moreover, it is generally believed that hot water freezes more quickly than cold water in vessels, and the argument in support of this is advanced that contrary is excited by contrary, just like enemies meeting each other. But it is certain that cold water freezes more quickly for anyone who makes the experiment . . . If hot water and cold water are placed in two vessels, the cold will freeze more quickly. Therefore, all things must be verified by experiment.

[12] How could you design an experiment to test this? [13] Can you be sure you know when a sample of water is *completely* frozen? [14] If the experiment gives such conclusive evidence, why does the question persist after all this time? We owe much to Roger Bacon for his insistence that we go to nature for the answers to our questions. However, you may be surprised at your results if you do this particular experiment.

On a cold winter morning one of the authors put outdoors three similar containers with equal volumes of water and a thermometer in each. The first contained cold water drawn from the hot-water tap before the pipe had heated. The second contained hot water from the hot-water tap. The third contained cold water from the cold-water tap. Table 5–2 shows the time and temperature readings as the water cooled and finally formed ice. [15] Do the results support Roger Bacon's assumption?

On the next day, when the temperature was a few degrees higher, the experiment was repeated. This time none of the containers filled with crystals. Each of them, however, had gone below zero. When a tiny crystal of frost was dropped in each, they filled with crystals in less than one minute and the temperature rose, in every container, to zero. [16] Why



Time sequence of a water sample as it cooled and froze.

Table 5-2

January, 1967
Air temperature −10°C
Observed temperatures

Crystals in both the hot and cold water from the hot tap appeared finer than those from the cold tap. None of samples were frozen solid; all three were a mushy slush.

Time	5:20 a.m.	5:50 a.m.	6:15 a.m.	6:30 a.m.	6:45 a.m.	7:05 a.m.	7:35 a.m.
Cold water from hot tap	14.4°C	5.0°C	0.0°C	−1.0°C	−1.8°C	0.0°C	−0.5°C
Hot water from hot tap	30.0°C	15.0°C	4.4°C	1.2°C	0.5°C	−1.4°C	0.0°C
Cold water from cold tap	14.4°C	5.0°C	0.0°C	−0.5°C	0.0°C	0.0°C	0.0°C
Additional observations		Bubbles in cold water from cold tap			Cold water from cold tap filled with crystals	Cold water from hot tap filled with crystals	Hot water from hot tap filled with crystals

did the results of this experiment differ from those of the day before? The situation is somewhat similar to what happened when you observed the supersaturated hypo solution in Chapter 3. [17] If this experiment is performed in an artificial cold environment, such as a freezer, do you think that a disturbance, such as opening the freezer door, would influence the results?

5–6 CONFUSION AND ORDER IN NATURE

Why does it seem to be so much easier to melt ice into water than it is to get an orderly crystalline arrangement of the ice again? It is because temperatures in our environment are more often above the point where water freezes, and because the crystalline pattern in ice is an orderly arrangement, whereas in the liquid state there is no evidence of order. Something about heat disrupts the orderly arrangement of solids and results in a condition of confusion. Depending on the kind of substance, the temperature at which this disorder becomes noticeable is different. This is another way of saying that different substances have different melting points.

Human life exists in a narrow temperature band where water is most often a liquid. We might expect the degree of disorder represented by the liquid form of water to occur in the rest of the universe. Similarly, if we lived in an environment where everything was "frozen"—where everything had an orderly pattern—we might expect the rest of the universe to be very orderly. Actually, on the sun, all matter is in the extremely disordered state we call *plasma*. Scientists believe that, as the intense heat of the sun and other stars is distributed among all the other matter of the universe, the total level of disorder is increased. It appears that the more ordered regions such as the earth lose some of their order.

Consider, at least briefly, man's part in the universe. At present his influence does not extend far beyond the surface of the earth, but he has produced some changes in his small region of influence which might be considered to have increased order. Man has taken materials from many different places and constructed buildings, roads, bridges, vehicles, and other goods of modern life. On the other hand, in using fuel, man takes materials such as coal and oil and breaks down order that has been built up over millions of years.

There is another aspect to the order concept as it relates to man—in fact to all living things. One of the characteristics of living things is that they can use disorganized matter to produce the tremendously complex and marvelously ordered substances of which life is made. All living

things are well-organized combinations of the common materials of the environment—soil, water, and air—and to these they will ultimately return when life is gone from their structure.

[18] Why do crystals form, then, if disorder is so common? We cannot credit inanimate particles of matter with the will to put themselves together. In the absence of excessive heat, however, these tiny bits of matter do assume a definite and characteristic arrangement. A valid theory of heat has to account for the fact that it will not only change the state of matter but that it will overcome the tendency to form regular patterns in certain materials. If heat is a fluid, it must be able to "dissolve the glue" that holds our particles together. But it cannot be a simple solvent, for, you will recall, when the ice crystals formed rapidly in the three jars, the temperature rose to 0°C. It was as if some of the caloric fluid was mysteriously released; heat appeared from within the water sample. The hypo, too, you will recall, was a source of heat when it suddenly started to crystallize. It seemed that some of the caloric fluid was available in excess when the more orderly system was established. We should add these observations to our theory of heat and conduct some special tests to try to determine the manner in which heat can contribute to the chaos of matter.

5–7 MEASURING HEAT

We have worked with quantities of heat but have not actually measured them. If the temperature change of a specimen depends on the amount of heat actually added, we might expect that it would also depend on the volume or mass of the specimen. We can predict that it would require twice as much heat to change the temperature of two equal volumes of water, for example, as would be required to make the same temperature change in one of the same volumes.

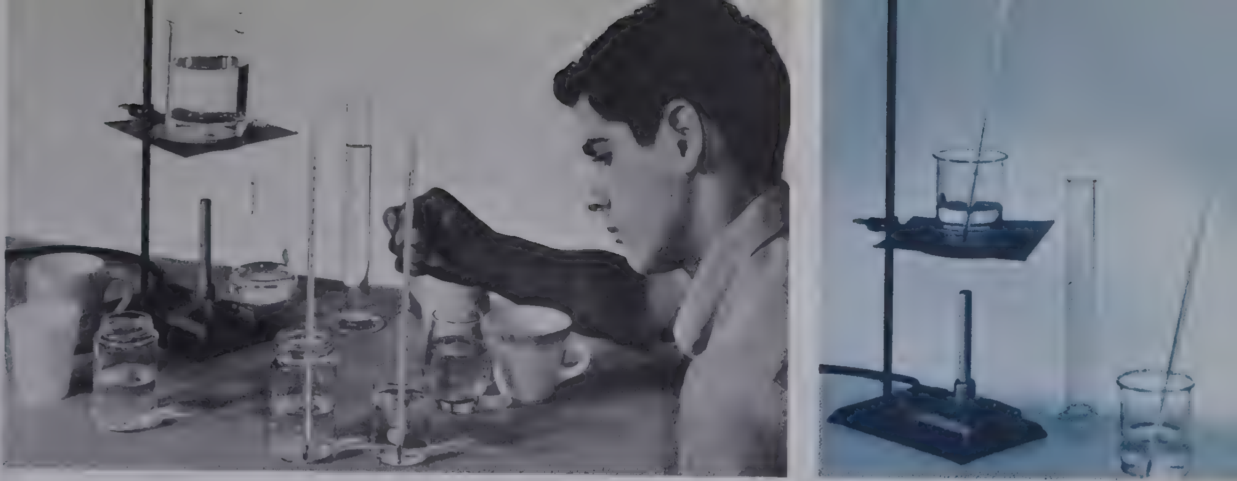
Let's check our prediction. Use the same materials as for Experiment 5–4, replacing the ice cubes with measured volumes of water. First heat 50 ml of tap water almost to the boiling point; record and plot temperature and time on a graph. Repeat with 100 ml of water and put the data on the same graph. [19] What would happen if you used 200 ml of water?

One factor that could cause these results to be slightly different from your prediction is that the can is also heated by the flame. Since containers must be used to hold liquids like water, it would be useful for us to determine the properties of different containers and their possible influence on our study of heat exchange.

5 6 You will need containers of similar size and shape made of various materials such as metals, ceramics, solid and expanded plastics, and waterproofed wood and paper; boiling water; and a thermometer for each container. Fill each container to the same level with boiling water, stir gently with the thermometer, and record the water temperature of each container at 30-second intervals over a period of 30 minutes. Plot a graph of temperature as a function of time for each of these containers, using the same sheet of graph paper for all containers. This will make it easier to compare the properties of the containers. Record any observations you make about the apparent temperature of each of the containers as you touch the outside at intervals during the test. [20] Does your recorded initial temperature for each container appear to be the same? [21] If they were each filled with water of the same temperature, how can you account for any differences in this initial temperature?

Our theory of heat must make provision for the observation that heat seems to move with different amounts of freedom through different materials. [22] If heat is indeed a fluid, can you explain why certain kinds of containers are more apt to allow this fluid to pass through them? Study those materials that are particularly good *insulators* (nonconductors) of heat and those that are especially good *conductors* of heat. Use a microscope if you think it will help. Try to determine the general characteristics that contribute to good insulation and to good conduction of heat. [23] If you have two containers that are different in color only, does this difference have an effect on the rate at which heat is lost? [24] If the fluid passes through the container walls, where does it go? [25] Is there any evidence that heat is lost from a container with an open top? If there is, how could this loss be reduced? [26] Does the thermometer itself absorb any of the heat? How could you determine this? [27] Is the temperature the same everywhere in a sample at the same moment? [28] Does stirring the liquid change the cooling rate?

5 7 We might compare the heat loss from our containers with the liquid loss from a container by leaking. For a simple experiment you will need a milk carton, a large needle or a small nail, a metric ruler, and water. Punch a hole in the bottom or near the bottom of the carton. Put the carton on a stand that will permit the water that leaks from the hole to fall into a sink or other large container. Stand a metric ruler on its zero end against the carton and tape it to the side so it does not move. This scale will read the depth of water in the carton. Now fill the carton with cold water and take readings of the water depth in the carton every thirty seconds. On graph paper, plot these depths vertically as a function of time and compare the shape of the curve connecting the points with the temperature curves you have obtained in Experiment 5–6. [29] What



Do all containers allow water samples to cool at the same rate?

influence would it have on your results if the hole were either larger or smaller? This experiment has shown differences in water loss through an opening in a container. [30] How does this help us to understand heat loss through different kinds of containers?

The kind of container that seems to show the slowest rate of heat loss will be the best for our continuing study of heat. We call such a container used for heat-measurement experiments a *calorimeter*. [31] Which one of the containers you used in Experiment 5–6 would make the best calorimeter? Why?

In the experiment with the simple thermometer made from glass tubing, we found that adding heat at a uniform rate caused the water to expand at a uniform rate. [32] Did the temperature also increase uniformly? We can use a different kind of experiment to find out if temperature is truly proportional to the heat added. We will mix two quantities of water that are at different temperatures and in this way add together both the total amount of water and the total amount of heat of the two.

5
8

You will need two calorimeters, two thermometers, a graduated cylinder, and hot and cold water. Measure 100 ml of cold water into one calorimeter, insert a thermometer, and record the temperature. In the other container put 100 ml of hot water and determine its temperature. Now pour the water from one container to the other and measure the final temperature after stirring the mixture. [33] How does this final temperature compare with the initial temperatures of the two parts? Repeat the experiment enough times to be confident that you can make a general statement about the temperature you would expect from such mixing.

Try to predict the temperature of the mixture of 50 ml of cold water at a known temperature with 100 ml of hot water at a known temperature.

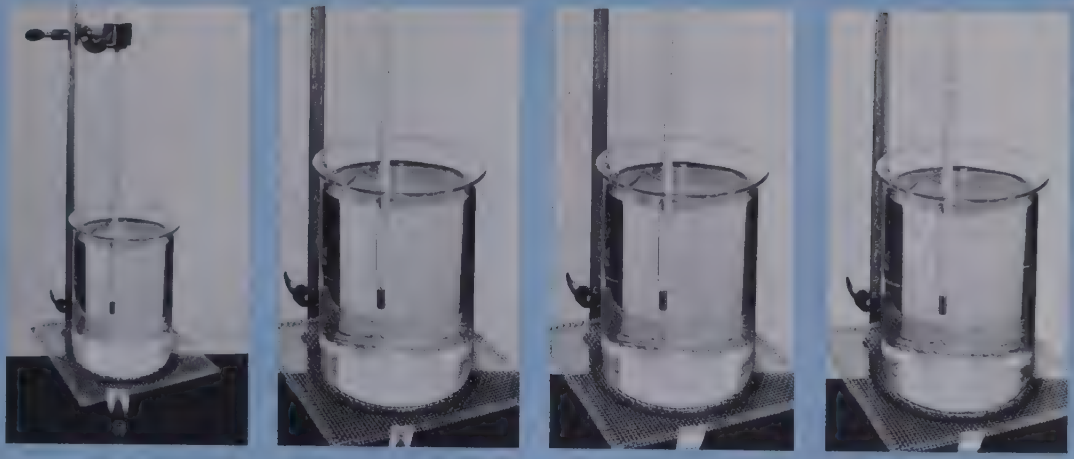
Then perform the experiment to check your prediction. If your prediction is sustained by experimental evidence, you have taken the first step in inductively arriving at a very important generalization: the principle of the conservation of heat. [34] Do the experimental results seem to be influenced if the hot water is added to the cold or vice versa?

Temperature may be observed in degrees on a thermometer, but the measurement of heat involves both the amount of the substance being heated and the temperature change it undergoes. We therefore measure it indirectly by bringing together these two quantities. We could invent our own unit for heat, but there is one that fits nicely with the metric units we have been using. It is called the *Calorie* and is the amount of heat required to change the temperature of one liter of water one degree Celsius. One calorie (with a small c) is the amount of heat required to change the temperature of one milliliter of water one degree Celsius. One Calorie (with a capital C) is equal to 1000 calories. Because we use two other measurements to compute it, the Calorie is a *derived unit*. Density is another derived unit that you have encountered. [35] From what two measured quantities is density derived?

If one liter of water is heated by 20 Celsius degrees, we have added 20 Calories to it. One-hundred milliliters (0.1 liters) of water cooling 40 degrees releases four Calories of caloric fluid. In each example the volume of water is multiplied by the change in temperature. Since you know the volume of hot water used in Experiment 5–8 and the temperature change that it underwent, you can calculate the number of Calories (or calories) that it gave up to the mixture. Similarly, you can determine the quantity of heat that the cold water absorbed. [36] Do you see now why it was so important to use a container with the best possible heat-holding ability? [37] Using the number of calories given up by the hot water, how could you compute the change in temperature that would occur in different quantities of colder water as it absorbed the heat lost by the hot water? (In these computations it will be more convenient for you to use the small calorie.)

Thus far, in talking about a single substance such as water, we have not been concerned about the quantity of heat the substance could hold without changing state. [38] What can you predict about the relative heat-holding ability of alcohol as compared with water? Remember that combining these two liquids resulted in a warming of the mixture. [39] How many calories of heat were made available when 100 ml each of the two room-temperature liquids were mixed? [40] Would you expect that dense materials would have more or less heat-holding ability than those that are less dense? [41] What other factors might be associated with heat-holding ability? [H, I, J]

What amount of heat in calories has the liter of water absorbed in five minutes? ten minutes? twenty minutes?



Zero minutes

Five minutes

Ten minutes

Twenty minutes

5-8 CALORIMETRY—THE METHOD OF MIXTURES

In Experiment 5-8 we suggested that you mix unequal quantities of water at different temperatures to add to your growing knowledge of the principles governing heat transfer. Let's continue to study heat transfer in water using a calorimeter.

Where we were mixing equal quantities of water at different temperatures, it was rather easy to predict that the final temperature of the mixture would be midway between the temperatures of the hot water and the cold water. If our experimental results supported this prediction, we might assume that there was a certain amount of heat contained in the total system. There was more heat in the hot water than in the cold, and some of the heat was used to raise the temperature of the cold water. In doing this the temperature of the warm water dropped. We can observe a similar situation if we connect two containers having different depths of water so they can flow together. The water flows from the deeper to the shallower level until both containers reach the same level. The total amount of water is still the same, although the distribution has changed. If these water containers were quart milk cartons with vertical sides, we could find the volume of water by multiplying the area of the bottom of the carton by the water depth in each. This volume would be the same before and after the mixing. What one side lost, the other gained.

How does the flow rate change as the water levels change?

Let's look at the heat exchange in two samples of water in the same way. The total heat may be related to the volume of water multiplied by the temperature. If we use the Celsius scale, temperature may be thought of as similar to the water level in each container and the freezing point of water can be thought of as the bottom of the container. We found in the cooling experiments that the rate of heat flow from a container was greater when the temperature was high than when it had cooled somewhat. Similarly, the rate of water flow is highest when the difference in water levels is greatest. Let's express the amount of heat in the hot water as the volume in milliliters of hot water multiplied by the temperature of hot water. By our definition of the calorie, the product is the number of calories of heat needed to warm this volume of water from 0°C to the hot-water temperature. The heat contained by the cold water may be found in the same way. The total heat is the sum of these two and should be the heat held by the mixture. Table 5–3 gives an example of some actual data taken in laboratory observations. Notice that the predicted temperature was within one degree of the observed temperature of the mixture. The difference might be explained by certain errors in measurement, so that a procedure such as this should be repeated several times to determine whether any differences are characteristic of

Observations

Trial	Hot water		Cold water		Mixture	
	Volume	Temp.	Volume	Temp.	Volume	Temp.
	a	*b*	c	*d*	e	*f*
1	100 ml	81°C	100 ml	10°C	200 ml	44.5°C
2						
3						

Calculations

Trial	Heat of hot water	Heat of cold water	Total heat	Predicted final temperature
	$a \times b$	$c \times d$	$(a \times b) + (c \times d)$	$\frac{(a \times b) + (c \times d)}{e}$
1	8100 cal	1000 cal	9100 cal	45.5°C
2				
3				

Table 5-3 Use the letters above each item of data as a guide to the calculations.

the method of performing the experiment. Notice also that the data taken do not include information about whether the hot water was poured into the cold or the cold was added to the hot. [42] Would you expect this to influence the results?

If the experimental results are so consistently near the predicted values that you believe that the method is valid, try some experiments using unequal quantities of hot and cold water. In the analogy with water volumes, you had water at two different levels in containers that were not the same size. If you actually experiment with water volumes, you might try a quart milk carton and a half-gallon carton. The final level of the two when connected or mixed will be more nearly that of the larger volume. Thus, by analogy, we might predict that if two volumes of hot water are mixed with one volume of cold, the final temperature will be nearer the hot-water temperature.

It is possible to measure the quantities of water used in the heat experiments more accurately by weight than by volume. [43] Why do you think this is true? We could determine the volume of different weights by using our knowledge that the density of water is 1 g/ml. You should do such an experiment for practice in weighing and because it is more suited to some of our later studies with solid objects.

Observations

Trial	Calorimeter wt.	Hot water + calorimeter	Hot water temp.	Mixture + calorimeter	Mixture temp.	Cold water temp.
	a	b	c	d	e	f
1	30.7 g	124.1 g	70.8°C	242.0 g	38.2°C	12.0°C
2	30.7 g	127.4 g	77.0°C	208.8 g	47.2°C	11.6°C
3						
4						

Calculations

Trial	Wt. of hot water	Heat of hot water	Wt. of cold water	Heat of cold water	Total heat	Wt. of mixture	Predicted final temp.
	$b - a$	$(b - a)c$	$d - b$	$(d - b)f$	$(b - a)c + (d - b)f$	$d - a$	$\frac{(b - a)c + (d - b)f}{d - a}$
1	93.4 g	6613 cal	117.9 g	1415 cal	8028 cal	211.3 g	38.0°C
2	96.7 g	7446 cal	81.4 g	944 cal	8390 cal	178.1 g	47.1°C
3							
4							

Table 5–4 If heat is conserved as it is transferred, an experimentally determined temperature should agree with the prediction.

5 9

You will need two calorimeters, two thermometers, a graduated cylinder, hot and cold water, and a balance. Weigh one empty calorimeter with its lid and thermometer and record this as the weight of your calorimeter (*a*). Add hot water to this calorimeter until it is about half filled and weigh it again. Record this value as "hot water plus calorimeter" (*b*). Also measure and record the hot-water temperature (*c*) just before mixing. Pour in cold water of known temperature (*f*) until the calorimeter is nearly full and take the final temperature (*e*) as soon as possible after mixing and stirring. Also weigh the mixture and calorimeter (*d*). It may be well to let your cold water stand briefly in another calorimeter to be sure you have the correct temperature. Record and calculate your findings in a manner such as is shown in Table 5–4.

All such experiments should be repeated several times to be sure that no important error such as an inaccurate temperature or weight

Observations

Trial	Calorimeter wt.	Hot water + calorimeter wt.	Hot water temp.	Mixture and calorimeter wt.	Mixture temp.
	a	*b*	*c*	*d*	*e*
1	32.5 g	181.0 g	59.7°C	219.3 g	31.5°C
2	32.5 g	155.9 g	77.0°C	199.8 g	36.4°C
3					
4					

Calculations

Trial	Wt. of hot water	Heat of hot water	Wt. of cold water	Wt. of mixture	Heat of mixture	Heat lost $(G - H)$	Heat to melt 1 g of ice
	$b - a$	$(b - a)c = G$	$d - b$	$d - a$	$(d - a)e = H$	$(b - a)c - (d - a)e$	$\frac{(b - a)c - (d - a)e}{d - b}$
1	148.5 g	8865 cal	38.3 g	186.8 g	5884 cal	2981 cal	77.8 $\frac{\text{cal}}{\text{g}}$
2	123.4 g	9502 cal	43.9 g	167.3 g	6090 cal	3412 cal	77.7
3							
4							

Table 5–5 More heat is lost by the hot water than we can account for by the warming of the cold water only. This heat is needed to melt the ice.

reading has appeared in the measurements. You should also compare your results with those of your classmates to be sure that a consistent error associated with your apparatus or your method of measurement is not affecting your results.

5 10

In Experiment 5–4 we observed temperature changes as ice melted, the water heated to the boiling point, and finally evaporated. We could use the data to determine the heat required to melt ice. Now, however, we may use the added method of heat measurement we are developing here to determine this quantity. Ice cubes straight from the freezer first warm to the melting point from the freezer temperature. If we mix such melting ice with hot water and find our final temperature, we have all the information we need. Table 5–5 shows the data obtained in such an experiment. Quantity (*G*) in this table, the heat of the hot water, is the total heat that the hot water could give up in cooling to 0°C. We could use any other temperature for our reference, but the freezing point of water (0°C) is very convenient. It is also the temperature of the water formed by the melting of ice. Thus, all the heat in the system is contributed by the hot water, and the heat of the hot water is the total heat of the system.

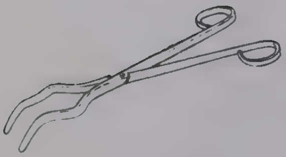
Notice the value we obtain in the calculation listed in the final column of Table 5–5. [44] How does this value compare with the amount of heat that raises the temperature of water 100 degrees to the boiling point? The value in the final column is called the *latent heat of fusion of ice*. The word *latent* indicates that there is heat present in water that is not readily detectable. This latent heat is the amount required to change one gram of ice to one gram of water at 0°C without a temperature change. The heat seems to disappear, but it reappears when water is frozen back to ice. So we must assume that it exists in some form within the water.

In theory we might also determine the heat released by steam (water in the gaseous state) as it condenses into water. That is, we might determine the latent heat of steam. Since the volume of a gas is determined by factors we have not as yet fully considered, it is not possible for us to measure the volume of the gas used, but only its volume of water after condensing. [K]

5–9 HEAT-HOLDING ABILITY OF SOLIDS

Metals are generally good conductors of heat and should exchange heat with their surroundings almost as readily as two liquids mix and come to the same temperature. Since metals are denser than water, a study of heat and metals may give us some insight into the relationship between density and heat retention.

5 11 You will need large samples of aluminum, iron, lead, and copper; a calorimeter large enough to hold each sample; two thermometers; cold water; a container of boiling water in which to heat the metal specimens; and a forceps to handle them. Measure enough cold water into the calorimeter so that the first metal sample will be covered but none of the water will spill out. By displacement of water, or by some other means, find the volume of the metal. With forceps or string, lower the metal object into the boiling water and allow the water to return to boiling. Be sure to record this temperature. Now transfer the hot metal sample to the calorimeter of cold water; stir and record the final temperature of the combination. From our knowledge of the temperature change in the cold water, we can calculate the number of calories of heat the water gained. From the principle of conservation of heat, we would expect that this number of calories is the same amount of heat given up by the



Determining the heat-holding ability of metals.

metal object. If we use our measurements of the volume of the metal and its change of temperature, we can determine the heat-holding ability in calories per milliliter of metal.

Follow the same procedure to determine the heat-holding ability of the other metals selected. You might also try cold metals in hot water, to compare the heat capacities of these substances when they are being warmed and when they are being cooled.

Perhaps we are not using the soundest judgment when we try to measure the heat capacities of solids in terms of the volume, because the volume of a sample is not constant. It may be that heat capacity is more closely related to the mass. Since the mass of one liter of water is one kilogram, we can as well define the large Calorie as the quantity of heat required to raise the temperature of one kilogram of water one degree Celsius. Likewise, we can define the small calorie as the quantity of heat required to raise the temperature of one gram of water one degree Celsius.

By weighing the metal samples used earlier, determine the heat required in calories per gram to raise the temperature of each kind of metal one degree Celsius. When we compare this number to the value assigned for water (1 cal per g per °C) we have the *specific heat* of the substance. Notice that we use the word *specific* to describe the relationship between some property of a substance when compared to that same property of water, as in the term *specific gravity*.

TAKING INVENTORY

Temperature is a measure of the intensity of heat in any given object but not of the quantity. The Calorie is the unit used to express quantity of heat. A small object can be very hot and yet contain a small quantity of heat. A large object can be only a few degrees above room temperature and yet contain a large amount of heat.

This chapter has stressed the idea that heat might be a fluid that is easily transferable from one object to another. According to this theory, heat is a substance. With the fluid theory in mind, we investigated the increase in volume with heating and the changes of state from solid to liquid to gas.

We also measured the exchange of heat between matter at different temperatures. The results suggest that heat, whatever it is, is conserved; that is, what one quantity loses, the other gains.

We found that some materials are good conductors of heat while others are good insulators. We also found that different materials have different heat-holding capacities.

Chapter 6 is a continuation of our study of heat. We will focus on the effects heat has on gases and whether a caloric fluid theory can explain these effects.

FURTHER EXPLORATIONS

A. The tape recording of the cricket chirping was made at 7½ inches per second. It was then played back at lower speeds to determine if the pitch of the chirps dropped along with rate. As the low-rate chirps were heard at reduced speed, certain distinctive sounds appeared. Is this a language of communication?

When several crickets are chirping, they seem to call and answer. By using a recording, could you make a cricket respond to the artificial signal? Could you change his chirp rate if you used a chirp recording at a speed a little different from that which you would normally expect for the temperature?

B. What problems would you encounter if you made a thermometer with water as the expanding liquid?

C. You may be familiar with the kind of thermometer in which expansion of solids gives a reading on a dial. Do these solids expand uniformly with the application of heat?

D. What is the physical meaning of the expression "twice as cold" in comparing the temperature of two objects? of the weather on two days? What would be the temperature on a summer day when it is "twice as hot" outside as the 20°C temperature in an air-conditioned room? On a warm day the Celsius temperature might be 27°C. What do you think the actual temperature might be on a winter morning when it was only "half as warm"?

E. We define the Celsius temperature scale in terms of the freezing and boiling points of water under certain special conditions. If you live where the atmospheric pressure is less than at sea level, you will find that you cannot raise the temperature of water to 100°C before boiling begins. Also, substances dissolved in water affect both the freezing and boiling temperatures. In Experiment 2–1 we melted ice and lowered its temperature by adding salt. What is the lowest Celsius temperature which may be obtained by adding salt to ice?

F. Prepare a solution of salt in water by adding 25 g of salt to 200 ml of water at room temperature. At what temperature will this solution begin to boil? Continue to evaporate the water slowly from this solution and record the temperature at equal time intervals. How does this compare with the temperature-time relationship when boiling pure water?

G. Examine a candy thermometer (the scale is probably calibrated in Fahrenheit degrees on which the normal boiling temperature of water is 212°). What is the meaning, in terms of what you have already learned about solutions and crystals, of the markings "hard crack" or "hard ball" at certain temperatures on a candy thermometer.

H. Prepare a saturated solution of sugar in water at about 100°C and determine its heat-holding ability (as compared to water of the same temperature) by pouring it into a known mass of cold water. Does sugar solution have a higher specific heat than water? Does salt change the specific heat of water in which it is dissolved?

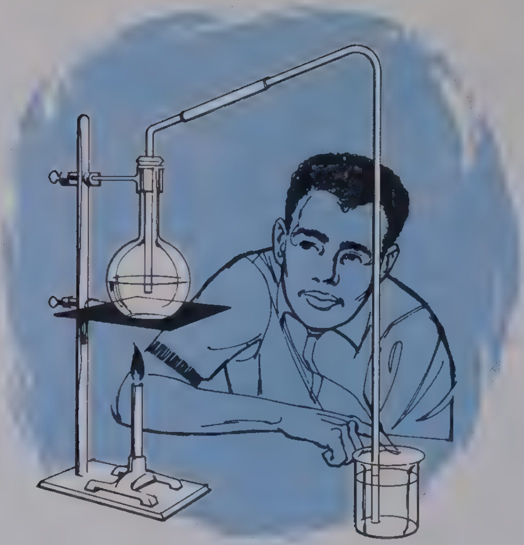
I. Ethylene glycol is the basic ingredient of some antifreezes. Conduct calorimetry tests to determine the specific heat of ethylene glycol. Determine its density and decide whether this substance is a better or poorer cooling agent than water.

J. Test different concentrations of several kinds of antifreeze to determine the relationship between freezing temperature and concentration. How do you determine when such a mixture is "frozen"? The reason we use antifreeze is to avoid damage to the automobile cooling system caused by the expansion of water as it freezes. Do all of these materials expand as they solidify?

K. As water goes through the change from liquid to gas at its boiling point, does it absorb heat energy as it does when it changes from solid to liquid?

To find out you will need a flask with one-hole stopper, a 60-cm length of glass tubing 6 mm in diameter, a burner, a ring stand and ring, a calorimeter, a thermometer, and two asbestos pads. Arrange the apparatus as shown in the illustration.

As heated water reaches the boiling point, the temperature stops rising. Where is the heat going? Can we assume that the heat is causing vaporization?



If so, can the heat be recovered as the water condenses in the calorimeter? Will a rise in temperature occur in the water in the calorimeter? What data can you collect that will allow you to calculate the value for the heat gained and lost as water evaporates and condenses? If you wanted to do this experiment to determine the value for the heat of vaporization, would it help to wrap insulation around the tube that connects the flask to the calorimeter?

L. In Experiment 5–11 you used the loss of heat from a metal at a known temperature (boiling temperature of water) to determine the specific heat of the metal. Now that you know the specific heat of the metal, could you use the same experiment to determine the temperature of the metal when heated to temperatures much higher than the boiling point of water?

FURTHER READING

ADLER, IRVING, *Hot and Cold.* The John Day Company, Inc., New York, 1959.
Describes experiments, measuring devices, theories of hot and cold, and temperature-dependent phenomena. Why ponds freeze from the top downward, why electric currents produce heat, and how to produce very low temperatures are among the topics discussed in this book.

LEERBURGER, BENEDICT A., Jr., *Josiah Gibbs.* Franklin Watts, Inc., New York, 1963.
A biography of one of the first American scientists. His pioneering efforts in the study of heat are described and explained in this book.

ROLLER, DUANE, editor, *The Early Development of the Concepts of Temperature and Heat: The Rise and Decline of the Caloric Theory.* Harvard University Press, Cambridge, Mass., 1957.

RUCHLIS, HY, *The Wonder of Heat Energy.* Harper & Row, Publishers, New York, 1961.

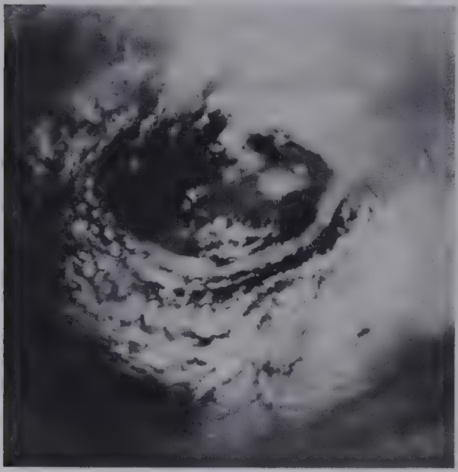
A general coverage of heat. Includes heating, refrigeration, heat transfer, and effects on gases.

SOOTIN, HARRY, *Experiments with Heat.* W. W. Norton & Company, Inc., New York, 1964.

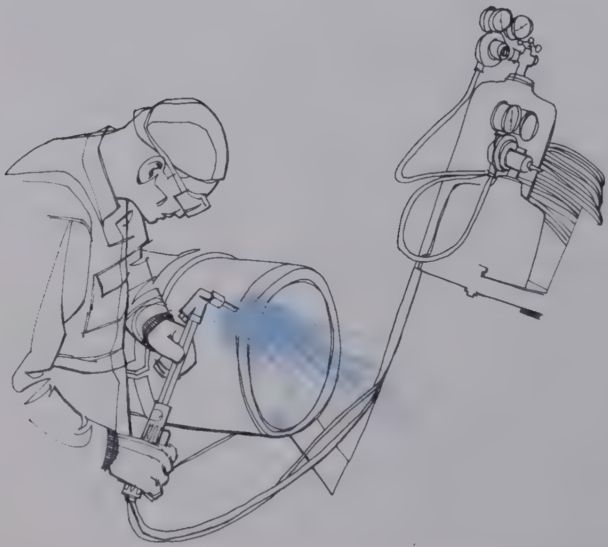
Contains many simple experiments with heat which can be done at home. Each experiment is accompanied by a summary of its scientific history and an explanation of the principles involved.

6

The Gaseous State



6-1 Gases and Heat
6-2 Testing for Caloric Fluid
6-3 Testing Volume Change
6-4 The Mixing of Gases
6-5 Another Motion of Gas Particles
6-6 Particle Motion and Temperature



A 250 - ml beaker of water will evaporate to fill the space outlined by the wooden frame.

6–1 GASES AND HEAT

Previously we have given our attention to the solid and liquid phases of matter and said little about gases. Because we are constantly surrounded by air, which is a mixture of several gases, we are inclined to take gases for granted. This state of matter is so important, however, that we need to study it carefully.

It is difficult to remove gases from any region. Liquids or solids may be added or removed from a container by simple processes like pouring or wiping, but this is not so with gases. We can pump air into a container and under certain conditions we can pump most of the air out of it. Regardless of which we do, the container is always full of air. Because of this, we cannot describe the quantity of a gas very effectively if we talk only of its volume.

Heat affects the volume of the gaseous state as it does the solid and liquid states. In order to study the relationship between gas volume and gas temperature, however, the pressure must be controlled, because pressure markedly affects the volume of a gas. It is important that we learn how to control gas pressure; in doing so, we will improve our understanding of both the structure of matter and the nature of heat. Since the air around us is the most readily available, we will use it for our study. To begin with, let's think about air at the pressure it is in the classroom.

Perhaps you already know that warm air rises above cold air. This motion, called *convection*, indicates that cold air is denser than warm air.

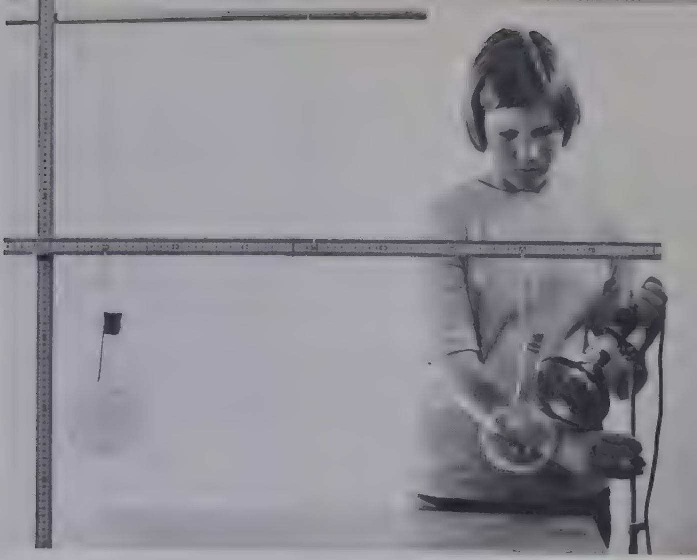
As the air is heated by the lamp bulb it rises and colder, heavier air moves in under it. The result is a convection current.

Recall that density is the ratio of mass to volume. If there is a smaller mass of air in a given volume, its density is reduced. Heating brings about this reduced density. What are the possible characteristics of heat that lower the density of a gas?

Any one of the following might be an explanation of the behavior of gases when heated. (However, they are not necessarily all of the reasonable explanations that could come out of our assumption that heat is a fluid.)

1. The caloric fluid which enters a sample of gas may have a negative mass. This would result in a reduction of the total mass. (The fact that we have not encountered negative mass previously does not mean that we can discard the idea without attempting to observe it.)
2. The caloric fluid provides additional volume, but does not associate itself with particles of matter. Then, the convection currents are only the results of this fluid spreading evenly throughout the gas particles.
3. The caloric fluid causes the particles of matter it encounters to repel one another more strongly than they did at lower temperatures. This implies that the fluid repels itself, but is attracted to particles of matter.
4. The caloric fluid increases the disorder of matter.

As we continue to think together along one line—that heat is a caloric fluid—keep in mind that there may be alternative explanations. Any



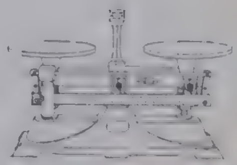
Will the addition of heat cause a change in the weight of the flask?

theory for heat that is consistent with all we have observed is potentially as good as those suggested here. You should try to design experiments that will test your ideas further and allow you to evaluate them against all other suggestions.

6–2 TESTING FOR CALORIC FLUID

One by one, let's put our ideas to the test of experimentation. Only one idea at a time should be tested. Then the answer obtained from the test can be associated with that idea with certainty. If several ideas are involved in the same test, the connections between ideas and answers are uncertain. For example, in an experiment to determine whether heat has negative mass, as in our first assumption above, we should remove the effect of any other influence. We cannot check this in an open container for we could not determine if there was a change in volume taking place.

For a test of whether heat has negative mass, you will need a two-pan balance, two similar flasks, two thermometers, two one-hole stoppers to fit the flasks, and a heat lamp. Carefully insert the thermometers in the stoppers; then put the stoppers on the flasks. With one flask on each side of the balance, adjust the slide or add masses to the balance until the system is balanced. Gently heat one of the two flasks with the heat lamp until its temperature is several degrees above that of the unheated one.



(**CAUTION:** Overheating may cause the stopper to blow out of the flask. Heat very carefully so this will not happen.) Remove the lamp and determine whether there is any evidence of a change in mass. Before we draw any definite conclusion, try to determine if any other conditions might have upset the balance.

[1] If you had used two unstoppered flasks, do you think the heating of one would have upset the balance? Try it and see. If there is a mass difference that might be due to the mass of the caloric fluid (either positive or negative), try to determine how great it is, and if it depends on the difference in temperature. [2] Do you believe that a change in mass alone could be responsible for convection currents?

The second possible characteristic of the caloric fluid that we listed earlier is contrary to our assumption in Chapter 5 that caloric fluid is attracted to matter. If caloric fluid does not associate itself in any way with the gas particles, we might never be able to detect its presence or describe it, for it is surely an invisible substance. [3] Is there some way we might test for this characteristic only?

The third suggested characteristic of the fluid is useful in explaining the expansion process, melting, and even boiling. A heated gas should readily work its way into a cooler gas of the same material. The fluid should distribute itself uniformly throughout the gas in a short time after the heat is removed. The temperature of the gas should soon become uniform and the circulating currents should stop. The caloric fluid can then be said to be uniformly distributed throughout the material. Heat would transfer only from a warm object to one of lower temperature because transfer in the opposite direction would be opposed by the repelling nature of the fluid for itself.

Since we cannot observe an orderly pattern for gas particles, we cannot test the fourth suggestion that the order might depend on the presence of the fluid.

6–3 TESTING VOLUME CHANGE

6 2 Let's design an experiment to check the relationship between the volume and the temperature of a sample of gas. Obtain or construct a gas expansion tube. It should be a glass tube about 20 cm in length, with the hole no more than 2 mm in diameter. Seal one end of the tube by heating and, while the tube and enclosed air are still warm, lower the open end of the tube into a droplet of mercury. As the glass and air cool, a bit of mercury will enter the tube and serve as a freely-moving piston to allow the enclosed gas to change in volume. (**CAUTION:** Mercury is a dangerous poison. Its vapor can destroy nerve and bone tissue if enough of it

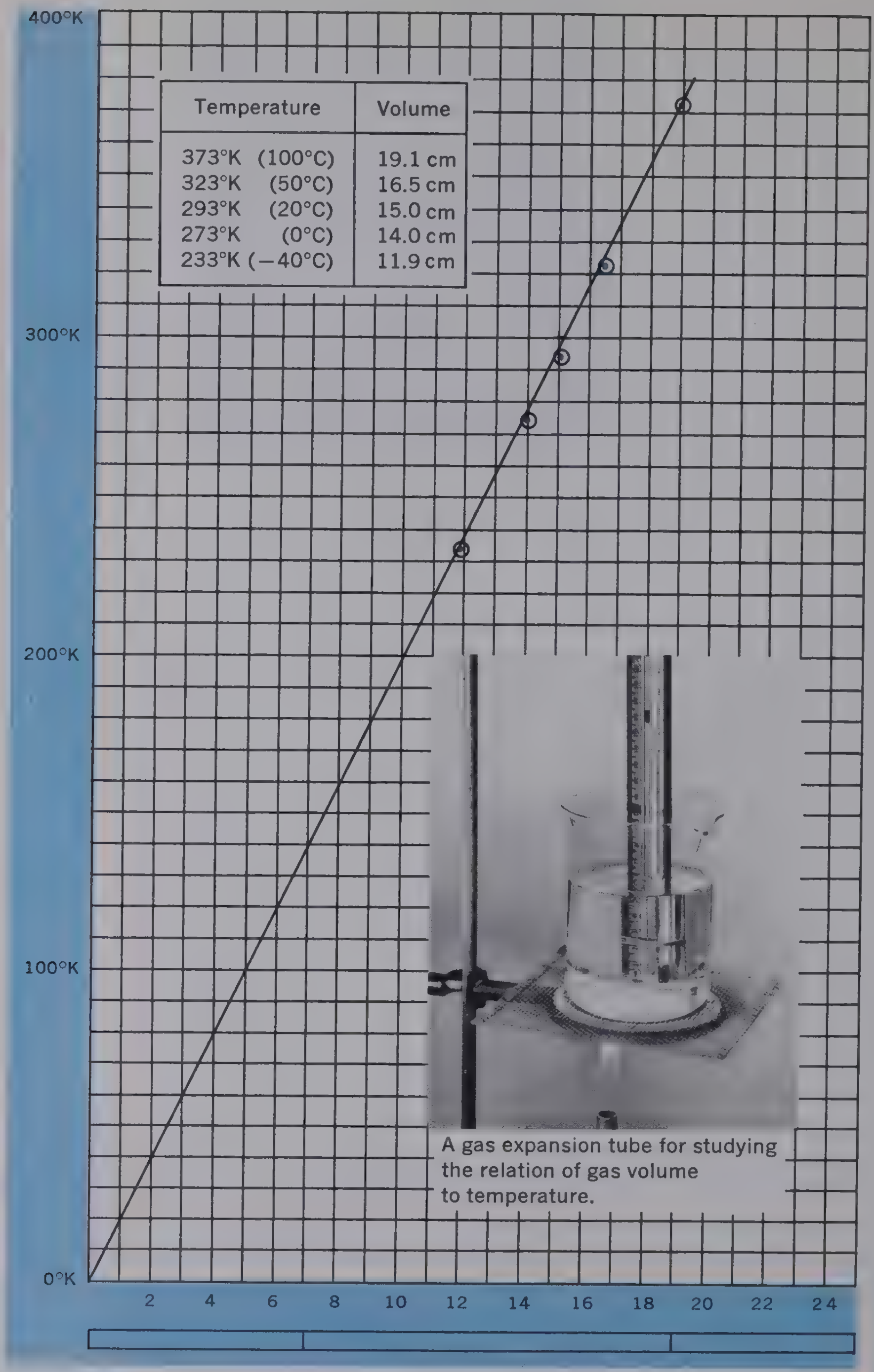
Sealing one end of a glass tube. The mercury will act as a piston in the cylinder inside the tube.

is breathed. Do not increase the vaporization of mercury by rubbing it on coins or jewelry or scattering it around. The apparatus suggested here exposes so little mercury to the air and at such reasonable temperatures that danger is at a minimum.)

Use rubber bands to hold the gas expansion tube, a narrow metric scale cut from a plastic ruler, and a thermometer together. Line up the closed end of the air column in the tube with the zero mark of the scale. Put this assembly in a test tube of ice water. Read the position of the bottom of the drop of mercury at the point of smallest volume and record along with the observed water temperature. Now gradually warm the test tube and make simultaneous readings of the temperature and gas column lengths every 5 to 10 degrees until the water reaches 60°C. Be sure the water is thoroughly mixed so that the temperature is uniform throughout at the moment when a set of readings is made.

Plot a graph of the column length as a function of temperature. As we noted in Experiment 5–3, when using a uniform cross section of tubing, the volume is proportional to the length. This graph is similar to a graph we would get by plotting volume of gas against temperature.

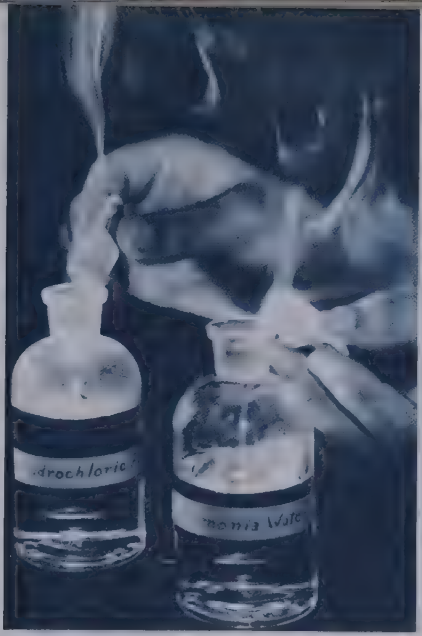
We can project our data by extending the graph beyond the lowest temperature and volume that we measured. At what temperature does the graph predict "zero volume" for the gas? This process of extending data is called *extrapolation*. It is important to realize, however, that we cannot accept predictions about what would happen in a region where we have not made measurements. The gas we examined was room air,



Temperature	Volume
373°K (100°C)	19.1 cm
323°K (50°C)	16.5 cm
293°K (20°C)	15.0 cm
273°K (0°C)	14.0 cm
233°K (−40°C)	11.9 cm

A gas expansion tube for studying the relation of gas volume to temperature.

Extrapolation of gas volume data to obtain a predicted "zero volume" for the gas sample.



When they come into contact, ammonia gas and hydrogen chloride react to form a white floating powder.

which is a mixture of several gases, principally nitrogen and oxygen. There is also some gaseous water in all room air. [4] Does the presence of this water vapor affect the rate of expansion? [5] Do all gases expand in a similar manner when subjected to the same change in temperature? [6] Does the thermometer measure the average temperature of the air sample?

The expansion of gases with increasing temperature indicates that, whatever the mass of heat may be, the change in volume could be responsible for some of the change in density. This does not show our second and third postulates to be wrong, but neither does it prove one to be better than the other. It has not clearly shown whether the gas itself, by some behavior pattern, increased the space it occupies or if the additional volume was occupied by heat. [A]

6–4 THE MIXING OF GASES

One need only walk into a room where some odor has been released to realize that gases travel through the air. We can detect when onions are being fried, when a cake is ready to come from the oven, or when toast is burned. Cologne released in one part of the room spreads to wherever we may be in the room. For all practical purposes these scents are at the same temperature as the room air, so for the most part this *diffusion*, as the process of mixing in liquids and gases is called, takes place without the addition of more heat. While we are aware that it is going on, it is a difficult process to study quantitatively, because there are already circulating air currents in most rooms. Besides, the major detector of gases in the air—the nose—is not designed to measure the amount of these gases but only their presence. It is worthwhile, however, to time the rate of diffusion of some pungent or aromatic gas from its source to observers at various places in a relatively still room. [7] In what way does one gas move through another? [8] Are the gases mixing, like the solutions of liquids and solids that we have studied before? [9] Or is one gas, flowing like a heavy oil, thrusting its way through the air?

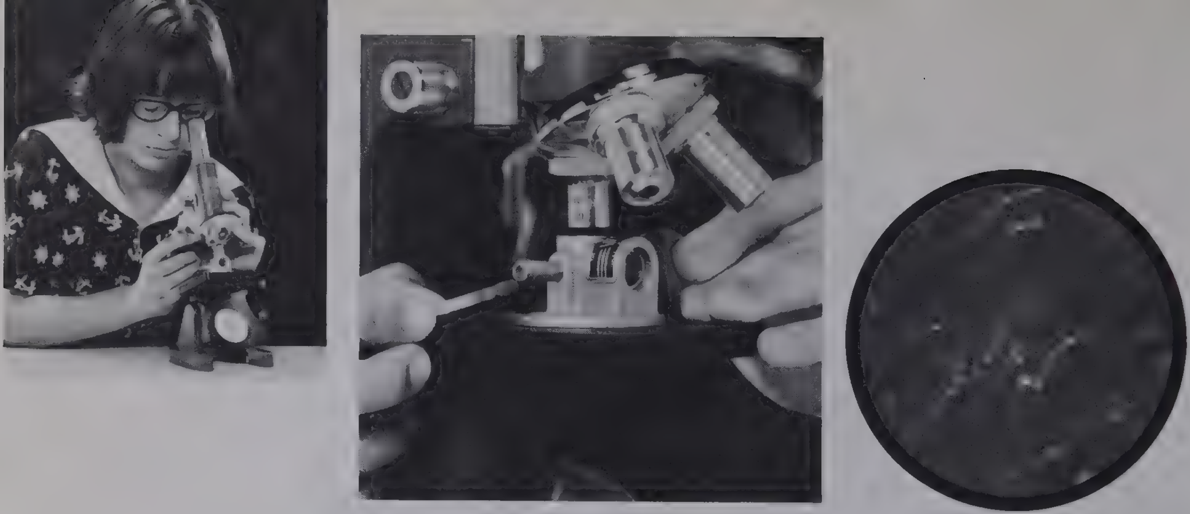
Hydrogen chloride and ammonia are colorless gases which react together to produce a solid substance. The presence of this new compound can be taken as an indication that the two gases have come in contact. Move the moistened stopper from a bottle of ammonia water near the open mouth of a bottle of concentrated hydrochloric acid. The white cloud which forms is a new compound, ammonium chloride. The ammonia gas permeates the air with its sharp odor. Hydrogen chloride, the vapor from the acid, also has a noticeable odor. The presence of the white precipitate and the odors is evidence that the ammonia and the hydrogen chloride are leaving the bottles and moving about in the air.

6 3 To study the diffusion of these two gases more carefully you will need a piece of glass tubing 2 cm in diameter and 50 or more cm long, two cork stoppers, two eyedroppers, and a thermometer. Fit the cork stoppers into the ends of the piece of glass tubing and place the tube horizontally on a table. Remove the stoppers and touch a drop of the hydrochloric acid to the small end of one stopper and a drop of ammonia water to the small end of the other. Put the stoppers in the ends of the tube simultaneously and record the time. When the diffusing gases meet they will form a ring of white substance—ammonium chloride. Determine the time required for these gases to make contact somewhere in the length of the tube. Measure the distance from each stopper to the ring of ammonium chloride; record this information along with the air temperature in the area where the experiment is conducted. Clean and dry the tube and repeat the process enough times to obtain a significant average for the contact time and for the distances between the stoppers and the ammonium chloride ring. [10] Are the time and the distances in agreement with those obtained when the experiment is performed with the tube in a vertical position, with first one and then the other chemical at the top? [11] Does the diameter of the tube influence any of these factors? Put the open tube in a refrigerator and allow it and the air inside to cool to the surrounding temperature and then repeat the procedure. Try to find

Diffusion of a gas or gases through air in a tube can be studied with this simple apparatus.

other environments at different temperatures and see if there is a relationship between the diffusion rate and the temperature. Both of these gases change the color of such chemical indicators as moist litmus paper. By introducing an indicator at one end of the tube, the motion of a single gas traveling from one end to the other may be studied.

That the odor from ammonia or some other source is not as strong throughout the room as it is near the source is an indication that the gas was diluted as it penetrated the air in the room. If we consider the ammonia gas in air to be a kind of solution, our particle model of matter would again be supported. Gas diffusing throughout the room indicates that such particles are naturally moving in some manner. [12] How can we determine what motion they have? Surely the particles of matter must be very small. They have defied any of our attempts to see them directly. If they may be condensed from gas form into liquids or solids we can see and measure, they must be very numerous. You have probably observed the motion of gases indirectly as the motion of dust particles or smoke suspended in the air in the path of a bright beam of light—the Tyndall effect—which we have discussed. Consider your observations and ask yourself how they fit into the caloric fluid theory and whether it can provide a model for gases as a good theory should. [13] If gas particles move about, is there another way, based on this motion, that can be used to explain our observations of heat? [B, C, D]

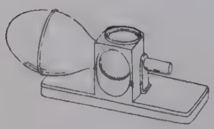


Motion of the smoke particles strongly suggests that other, invisible particles are buffeting them about.

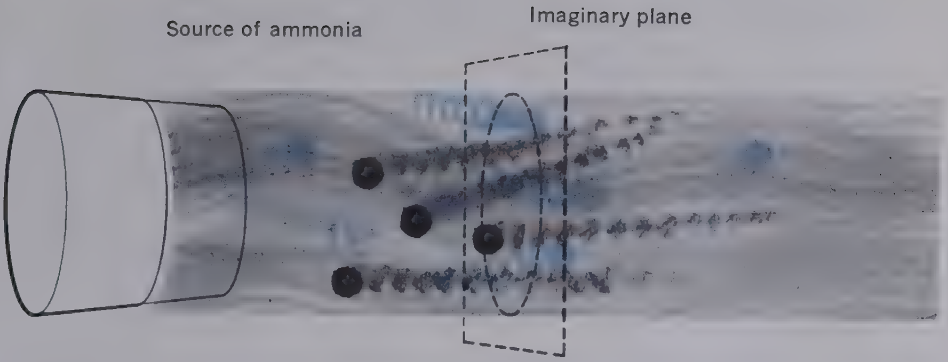
6–5 ANOTHER MOTION OF GAS PARTICLES

There is another, more subtle motion, especially of the smallest particles, in the gas currents. To observe it we must watch particles too small to see without magnification. Other causes of motion must be minimized and the magnification must be quite high.

You will need a microscope, a strong light source, and a small chamber to confine and protect the observed particles. The small chamber is available commercially or one may be made. It should be small enough (no more than 1 cm thick) to fit under the low power objective of your microscope and have a clear glass or plastic top to permit observation. At least one side of the chamber should be clear to admit a bright light to illuminate the contents and there must be an opening to admit smoke, which contains particles large enough to be seen with the aid of the microscope. Put the chamber on the microscope stage as shown in the illustration. Fill it with smoke and set up a bright light to illuminate it from the side. (Do not attempt to illuminate it from below the stage. In fact, it is best to have the bottom covered with black paint or tape.) Through the low power of the microscope you may see light reflected from the individual particles of smoke.



This motion of smoke particles does not seem to be caused by convection currents. Instead, it appears that the smoke particles are being buffeted about by some invisible agent. Perhaps the smaller and unseen



The movement of ammonia particles in the diffusion tube can be visualized by using an imaginary plane.

particles of matter are somehow knocking the larger particles about. [14] Is this caused by heat? Although the gases we are observing are at room temperature, they do have some heat content. If there were a way to study this chaotic motion at different temperatures we might be able to determine if it is related in any way to the heat that had been added. [15] Do our observations from the experiment with diffusion of the two gases in the tube help to explain what we have just seen?

If particles of air are moving fast enough to cause our smoke particles to dance so wildly, they, too, must be in very rapid motion. [16] Why, then, does it take a relatively long time for one gas to diffuse through another? [17] How can we reconcile the oil-like flowing of ammonia along the tube with the bulletlike speeds of the particles of the two gases?

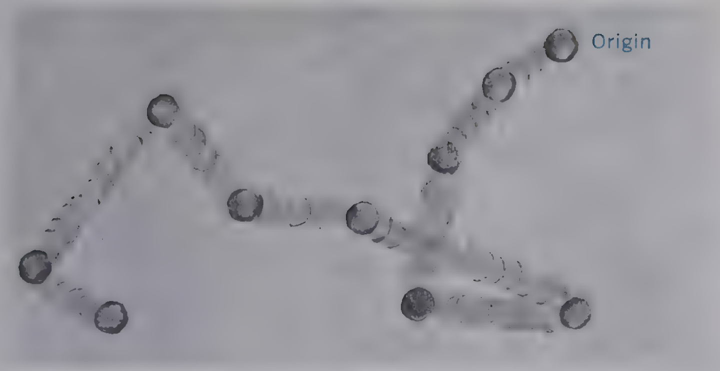
Let's consider the situation in the tube that enclosed the diffusing gases we studied earlier. Near one end of the tube there is a high concentration of ammonia gas particles. If the ammonia flowed like honey or oil, it would replace the air where it was entering. We know, however, that no air leaves the tube because of the stoppers. There must be a concentrated solution of ammonia in the same air that was in the tube. Look at the illustration above of the tube we used in Experiment 6–3. Imagine that we can study what is going on in a very narrow section of the tube; this section is represented in the drawing by the dotted oval where the imagined plane intersects the tube. On the left of this plane we will assume there is a high concentration of the ammonia gas in the air in the tube. On the right, farther from the source, the concentration of the gas in the air is lower. We will further assume that there is an aimless, random motion of the particles; as they travel down the tube they collide and rebound millions of times. There is no preferred direction for any one of them to move at any particular instant. Out of this random motion, however, some small fraction, say 1 percent of the ammonia particles just to the left of the imaginary plane, move over to the right of the imaginary plane in a certain time interval. Likewise, 1 percent of the ammonia particles just to the right of the plane should cross to the left side in the same time. But there were more ammonia particles on the left, near the source, than there were on the right. A greater number moved from left to right than from right to left. This would tend to make the concentrations more nearly equal. Eventually, as we have already considered in crystal growth, equilibrium is reached when equal numbers cross our imaginary plane in each direction during a given interval of time. Eventually the ammonia becomes uniformly mixed with the air in this region.

The region we were considering was imaginary, however, and could have been anywhere in the tube. We may then see how a theory of high speed particle motion can be used to explain how one gas becomes completely dispersed throughout another. We should try to decide if this motion might be caused by the presence of a caloric fluid which converted our substance to the gaseous phase, or if there is another possible explanation.

6–6 PARTICLE MOTION AND TEMPERATURE

How can we explain that the rate of diffusion in a gas changes with temperature change? First, we know that at higher temperatures gases are less dense; there are fewer particles per unit of volume. Second, if heat is associated with particle motion, the particles move faster at higher temperatures. Try to use these facts to build an answer to our original question.

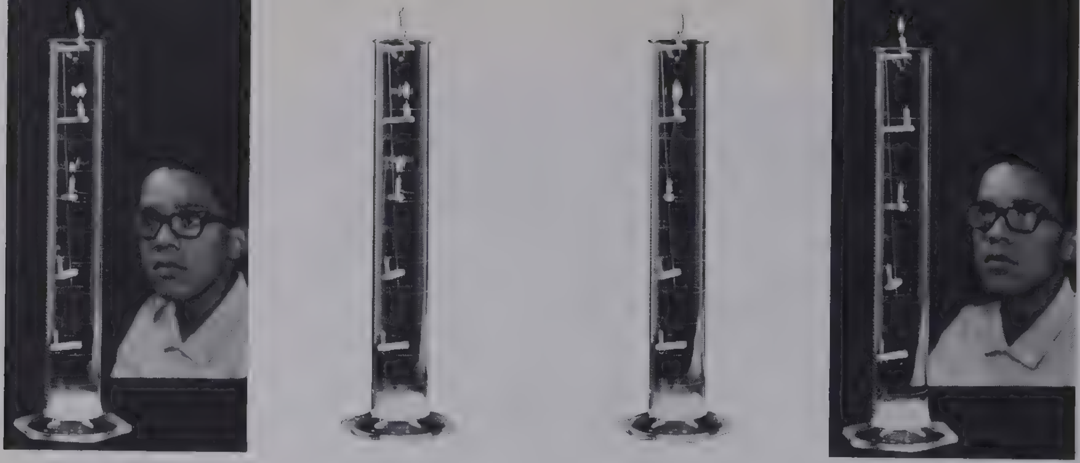
How much each of these two factors, density and particle speed, contributes to rate of diffusion is something we are unable to investigate because special equipment is needed. We would have to cool or heat a sample of gas without allowing a change in density. To do this we would need a very strong sealed container equipped with a valve system so gases could be introduced and removed. With such equipment we would be able to observe the effect of particle motion separate from the effect of density and know what each contributes to the rate of diffusion of gases.



A representation of the erratic path of a single gas particle.

[18] Can a theory that attributes motion to gas particles explain the expansion of a gas with heating that we have observed? A greater rate of motion of particles would result in an increase in the collision rate within the same space. To a certain extent, it would be like the conditions on a dance floor when the tempo of the music increases. As the dancers increase their activity, space needed by each couple increases or the collision rate increases. To avoid collisions some dancers move apart to provide more room. It is probably a coincidence that this occurs when the music gets "hot." Perhaps the gas particles, too, require more room as they increase in rate of motion and collide in some way with their neighbors with greater impact and greater frequency, unless the space is enlarged.

One more relationship between the condition of gases and their temperature comes to our attention if we think of how warm the hose of a tire pump gets when we inflate a tire. [19] Is this condition associated with pressure? When we release air from a tire, the stream of exhaust gas seems quite cool. Is it actually cool or is our nervous system reacting to increased evaporation from our skin? Place a thermometer bulb in a stream of air coming from a high pressure source and expanding to the pressure of the room. The gas coming from a carbon dioxide fire extinguisher is much colder than the container from which it came. [20] What happens to the temperature of the fire extinguisher itself after much gas has been released? Compare this with the temperature change of pressurized cans of spray paint or hair spray immediately after use.



To what height has the carbon dioxide risen?

Now let's look at some examples of gases in which diffusion does not seem to take place readily or is influenced by some external condition. Carbon dioxide may be obtained by allowing a piece of dry ice to evaporate in the bottom of a beaker. By thrusting a glowing splint into the top of the beaker occasionally, you may determine the time required for the beaker full of gas to diffuse through the open top of the container, because this gas will extinguish a spark. [21] Is this rate about the same as the rate at which you would expect any gas to have diffused under similar circumstances? [22] Do your data on the diffusion rate of hydrogen chloride and ammonia in the vertical tube in Experiment 6–3 indicate other factors than those we have discussed which may govern the rate of motion?

Consider the situation in the atmosphere where large bodies of cold air "front" upon warmer air masses. Such a cold front may be accompanied by temperature drops of thirty or more degrees per hour. Here, evidently, are quite different gases—at least they have decidedly different temperatures and densities—which do not appear to be mixing at a very rapid rate. [23] Does this support or contradict our other findings about the mixing of gases? Gases may also diffuse through seemingly solid materials. A toy balloon inflated with hydrogen will shrink in size in time, although no holes can be seen in the surface of the balloon. [24] Do air-filled toy balloons appear to shrink at the same rate? Do all gases escape at the same rate from balloons? Allow one inflated balloon to stand in the sunlight; compare its shrinking rate with a second inflated balloon kept in shade.

Gases must also diffuse through living tissue as part of the respiration of plants and animals. Certain kinds of tissue, such as the leaves of plants, the gills of fish, and the lungs of warm-blooded animals, are specifically adapted for this. In our lungs, inhaled air contributes oxygen to the blood stream while gaining carbon dioxide from the blood. Thus there is simultaneous transfer of these two gases in opposite directions.

In trying to check our second and third explanations in Section 6–1 (page 126) of the behavior of gases when heated, we have found reason to consider the fourth explanation. We began our latest experiments to test a caloric fluid that attached itself to matter and caused these particles of matter to repel each other. We have not contradicted such a concept, but to explain all we observed, we must recognize that the particles are also in a state of random motion.

One of the purposes of scientific theory is to unify many seemingly unrelated observations. If a theory gives a satisfactory explanation for many things and yet remains simple and uncomplicated, it is, indeed, a very good theory. If a theory cannot explain observations, then it becomes suspect. Since we cannot deny the apparent motion of the particles of matter after these experiments, we cannot accept a theory of heat that does not give some consideration to that motion. A caloric fluid theory is now not enough. It does not adequately explain all our observations. If the fluid exists, it acts along with the moving particles of matter. [25] Can we explain all we have observed on the basis of particle motion alone? [26] Is the idea of particle motion a basis for a better theory, one that will suggest new experiments? [E, F, G, H, I]

TAKING INVENTORY

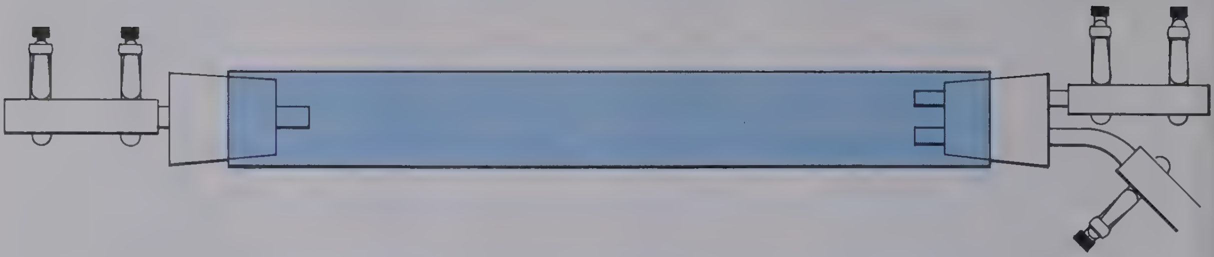
In this chapter we have examined the changes in density of gases with temperature, the effects of heat on the volume of gases, the mixing of gases by diffusion, and the motion of microscopic particles that mingle with gases. We have also examined the adequacy of the caloric fluid theory and the particle motion theory as explanations for the things we observed.

At this stage of investigation the idea of particle motion seems to be a necessary part of any theory of heat. In Chapter 7, we shall examine the two theories still further in our effort to refine them into a single theory.

FURTHER EXPLORATIONS

A. If gases expand uniformly we should be able to use this principle in making a thermometer. What problems would we encounter in putting this principle into use?

B. What effect on diffusion rate would occur if we reduced the amount of gas in a container without changing the temperature? According to our third hypothesis in Section 6–1 (page 126), we would be removing some caloric fluid along with the particles removed.



C. Observe the structure of the ring of ammonium chloride that formed in the tube used for the gas diffusion experiment. Record any changes in its appearance over a period of time and try to fit this into our growing theory of the structure of matter and the nature of energy.

D. Make a further study of the diffusion of gases by devising a method of observing the progress of ammonia or hydrogen chloride along a tube using pieces of litmus paper or other indicator paper.

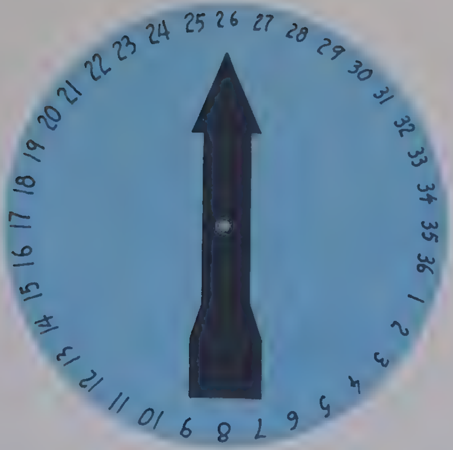
E. Set up the apparatus as shown in the illustration above. With a vacuum pump or aspirator, pump out the tube, then simultaneously release the inner clamps to release a drop of each liquid into the tube. Observe the nature and rate of diffusion.

F. We make reference to heating a tire pump and cooling a carbon dioxide fire extinguisher. Try to analyze these situations further. In one case, where is the heat coming from and in the other where is it going if cold is the removal of heat? In each, what is cooled and what is heated?

G. Build a mental model of what happens to air particles in a tire pump as the pump is used. Can your model explain why heat is evident as a result of pumping?

H. Build a mental model of what happens to air particles as you let them out of a tire. Can the model explain why an object placed in the stream of air is cooled? Would your model apply to carbon dioxide escaping from a fire extinguisher? Explain.

I. How might a model for the random path of a gas particle be established? One way to establish such a model is to plot the random numbers that are obtained from spinning an arrow on a numbered wheel. To make such a wheel, draw a circle 8 cm in diameter. Use a protractor to locate 10° intervals on its circumference. Number these intervals from one to thirty-six. Cut out an arrow 6 cm long and center it on the circle, then use a thumbtack or a pin to give it an axis of rotation.



Spin the arrow once; the interval times 10 will be the angle with the horizontal that the particle takes on the first leg of its path. Spin the arrow again; the interval, expressed in mm, can be the distance traveled on this path before the particle collides with another particle. A third spin of the arrow provides the angle of the old path with a new path and a fourth spin identifies the distance traveled before another collision occurs.

Plot the angles and path lengths for about a dozen collisions and you will begin to have a feeling for the random, completely haphazard motion that seems to be characteristic of gas particles.

FURTHER READING

CONANT, JAMES B., editor, *Robert Boyle's Experiments in Pneumatics*, Harvard University Press, Cambridge, Mass., 1950.
Traces the development of our understanding of the nature of gases. Includes original papers by Boyle plus a discussion of his experiments.

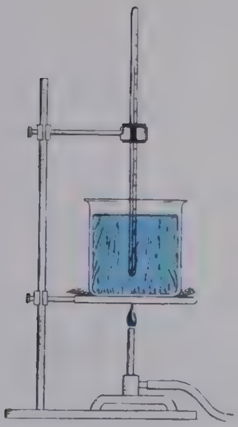
EFRON, ALEXANDER, *Liquids and Gases*. John H. Rider Publisher, Inc., New York, 1958.
A theoretical and practical treatment of liquids and gases.

LEHR, PAUL, and others, *Weather: A Guide to Phenomena and Forecasts*. Golden Press, Inc., 1957.
This handbook on weather discusses many principles that apply to a study of gases.

SHAMOS, MORRIS, editor, "Boyles Law: Pressure Volume Relations in a Gas—Robert Boyle," in: *Great Experiments in Physics*. Holt, Rinehart and Winston, Inc., New York, 1959.
Contains a brief biography of Boyle and gives a modernized version of his original paper describing the study of the relation of pressure and volume of a gas.

7 Looking Back: Reviewing the Evidence and the Theories

7–1 What We Know About Heat Theory
7–2 The Two Heat Theories and Evaporation
7–3 The Two Heat Theories and Boiling
7–4 The Two Heat Theories and Diffusion
7–5 The Two Heat Theories and Solution and Crystallization
7–6 The Two Heat Theories and Conduction



7–1 WHAT WE KNOW ABOUT HEAT THEORY

In the beginning of our studies, we proposed three theories for the structure of matter: It is continuous; it is composed of particles, or it is both. The evidence gathered during our work with solutions and crystals favored the particle theory. Our study of gases provided further support to the idea that matter seems to be made of tiny particles. We have also found that these particles may be closely packed together (as in solids), loosely organized (as in liquids), or separated still farther (as in gases).

Since changes of state seem to be related to the temperature of matter, we explored the topics of temperature and heat to determine what causes temperature to vary. We found that when heat is taken from matter, the temperature drops; when heat is added to matter, the temperature rises.

It is time to summarize the evidence we have collected and examine its implications to improve our theories of both the structure of matter and the nature of heat. We will concentrate on the two possible heat theories.

The Caloric Fluid Theory

The caloric fluid theory assumes that heat is a substance. This implies that it has mass and takes up space. Our investigations allow us to say the following about heat as a substance:

a. It has little or no mass.
b. It is colorless and cannot be seen.
c. It readily penetrates gases, liquids, and those solids we call conductors.

d. It does not enter readily into those solids we call insulators, or nonconductors.
e. It is released from some forms of matter during burning.
f. It is absorbed by some forms of matter during heating.
g. It appears to flow from regions of high concentration toward regions of low concentration (from hot objects to cooler ones).
h. It is involved in changes of state of matter.
i. It causes matter to expand.
j. It increases disorder in matter.
k. It has not been observed except in association with matter.

The Particle Motion Theory

The theory that heat is motion of particles of matter has also been considered. To use this explanation for our observations, we must assume that all matter is composed of particles, that these particles are in constant motion, and that the rate of particle motion increases as heat is added to matter and decreases as heat is taken from matter. Using the list of items that summarize our view of the caloric fluid theory, we can formulate a similar list of observations from the viewpoint of particle motion.

a. Rather than being a substance, heat is particle motion. It does not have mass or volume.
b. Particle motion (heat) need not have a property such as color.
c. Particle motion varies in solids, liquids, and gases.
d. If heat is particle motion, then it passes easily along conductors by collisions of particles.
e. Particle motion increases in burning.
f. Heat, in the form of particle motion, is exchanged by particles.
g. Hot particles, having the highest rate of motion, give some of their motion to cold particles. This cools the hot particles by slowing them down and warms the colder ones by speeding them up. Eventually the motion (or heat) becomes uniformly distributed.
h. The changes of state in matter indicate a change in the motion that particles can have.
i. Hot particles are moving at a higher rate than cold ones. They collide and bounce apart more violently, tending to separate and occupy more space.
j. The higher the rate of particle motion and the more random it is, the higher the disorder of the matter being heated.
k. Motion of particles cannot be isolated from matter.

Evaporation occurs at all temperatures. Heat taken from a swimmer's body as water evaporates makes him feel chilly.

On the basis of these lists, the two theories seem to stand about even. Either one could account for most of our observations. We should now look again at our earlier observations and continue investigating in order to select one as more probable than the other.

7–2 THE TWO HEAT THEORIES AND EVAPORATION

Let's look back at observations already made to seek a clearer understanding of the behavior of matter at different temperatures. During evaporation, tiny particles or a continuous stream of matter escape the body of liquid into the space above it. This occurs very readily at the boiling point of the liquid. But evaporation also occurs at lower temperatures. You know that wet clothing dries when hung on a clothesline and that water disappears from an open container. Standing in a wet bathing suit on a windy summer day, you have probably felt cold even though the air temperature was quite high. The heat necessary to cause the water particles to evaporate was taken from your body. Hence you soon felt chilled.

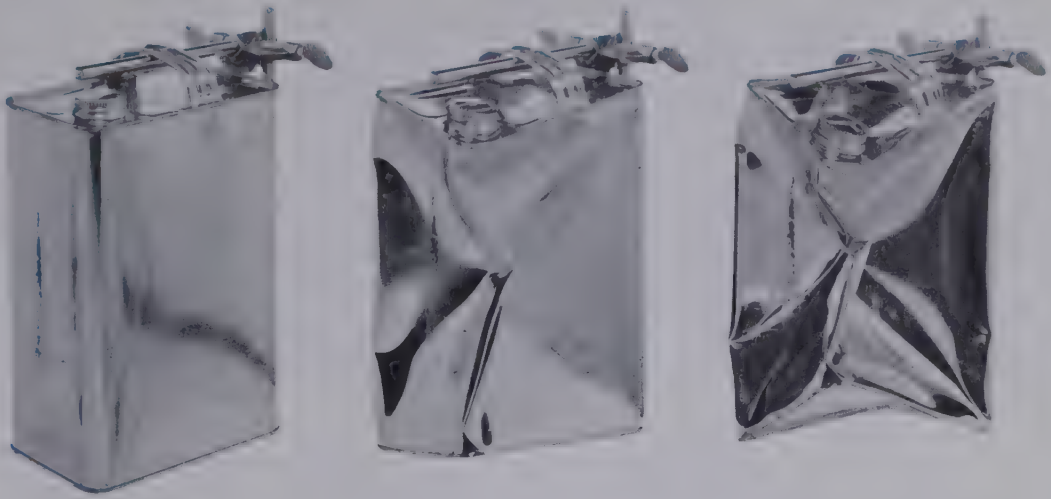
Which beaker of water will evaporate faster?

7 1 At temperatures below the boiling point, the evaporation rate is controlled by a variety of conditions. To learn about these conditions you will need a 250-ml beaker, a test tube, and a test tube rack. Put one 10-ml sample of water in the beaker and another 10-ml sample in the test tube. Set these two containers aside in the same part of the room. Check them regularly to determine the times required for each to evaporate completely.

[1] How would you use each of the two theories of heat to explain the difference in the rates of evaporation? Put two equal quantities of water in similar containers and locate one in an area sheltered from air currents while a fan blows across the other. Both should be at about the same temperature. [2] Again, attempt to use the two heat theories to explain your observations.

Not all the observations about the factors that seem to control evaporation indicate the merits of one theory over the other. We find that the surface area and the motion of the air affect evaporation rate, but this does not tell much about the nature of liquid as it evaporates. [3] Visualize what is happening to the water as it evaporates, first using the caloric fluid theory, then the particle motion theory.

Put a drop of alcohol on the back of each hand and blow gently on one of them. Record your observations and try to relate this to the



theories we have been testing. [4] Would drops of water have produced the same effect? [5] Does the fact that air from your breath is already warmed by the body influence the effect? There is usually more water vapor in exhaled breath than in room air. [6] Could this affect the results? [A]

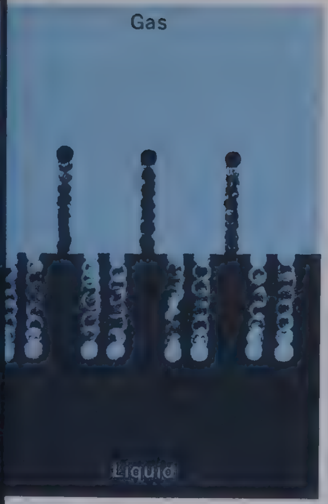
7–3 THE TWO HEAT THEORIES AND BOILING

To learn more about the nature of boiling, we can do a modification of a familiar demonstration. You have probably observed the effect of atmospheric pressure on a can when the pressure inside is sharply reduced. It is customary to boil a small amount of water in a gallon can and then remove the can from the heat, sealing it with a lid or stopper. As the outside of the can is cooled, it crumples from the tremendous forces resulting from the differences in pressure inside the can and outside. Usually this demonstration is done to show the great force of the weight of the atmosphere. We are more interested in observing the effect of reduced pressure inside the container on boiling. Let's do the same experiment, but instead of the destructible can, we will use a rigid glass container which can withstand the force.

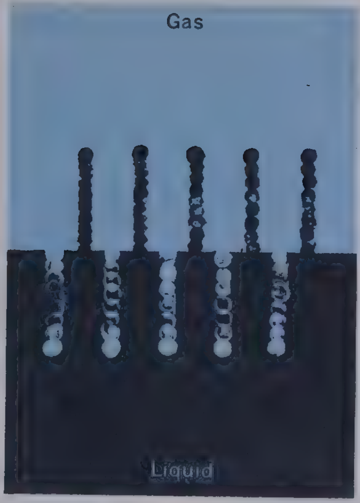
What is happening at the surface of the water in the sealed flask?

You will need a thermometer, a flask, a one-hole stopper to fit the flask, a source of heat, an asbestos pad, and a ring stand and ring. Place the thermometer through the one-hole stopper so that it reaches one fourth of the way to the bottom of the flask which the stopper will seal. Fill the flask one-third full of water and, with the mouth of the flask open, boil the water vigorously for two or three minutes. This will fill the flask with water vapor that contains very little air. Quickly insert the stopper combination into the mouth of the flask as you remove the flask from the heat source. Make sure the stopper is very tight, then invert the flask, placing the neck through a ring supported by a stand. Cool the flask above the water level with a wet cloth and observe both the water and its temperature. From the similarity to the experiment with the crushed can, we assume that the pressure inside the flask is very low when the flask is cooled. When we stop cooling the flask, the boiling rate subsides; then when it is cooled again the vigorous boiling resumes. [7] What can you predict about the relationship between the temperatures in the vapor and in the liquid in this experiment? [8] How could you verify your prediction?

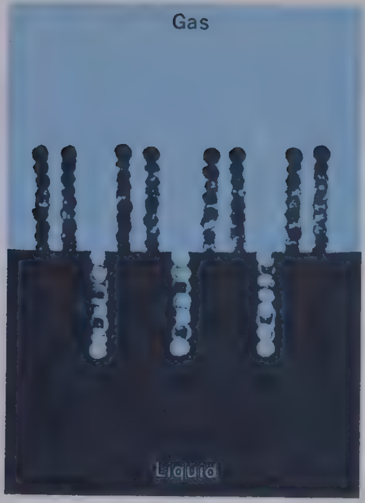
A condition of equilibrium exists between the liquid and the vapor phase of water during those times when we are not changing the heat of the system. It appears that the release of vapor during the boiling increases the pressure inside the flask and the system reaches the balanced condition. Cooling the upper part of the vessel condenses some of the vapor to liquid, reducing the pressure and upsetting the equilibrium. In an attempt to restore balance, some of the liquid changes to vapor and boiling occurs. Notice that, as the temperature of the water drops, it becomes more difficult to reach the boiling point. Finally, the water in the flask is so cool that further attempts to cool the upper part of the flask



evaporation < condensation



evaporation = condensation
(equilibrium)



evaporation > condensation

will not make it boil. This would lead to the assumption that liquid water and its vapor can be in equilibrium at any particular temperature. [9] Which theory of heat explains this behavior more satisfactorily?

We assumed that one of the properties of a caloric fluid is its tendency to cause the particles to which it attaches itself to repel one another. The act of cooling the flask removes some of this heat fluid; it also reduces the repulsion of the vapor particles so they can come much closer to one another. This allows the particles to condense. The heat fluid remaining in the liquid portion forces some of the water out of the liquid and into the gaseous phase. When the heat fluid is evenly distributed in the system there are no further changes that can be observed.

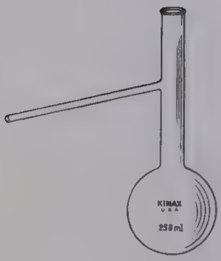
Using the particle motion theory, we assume that the particles of matter in the vapor are moving about more freely than those in the liquid phase at the same temperature. However, the gas particles have the same average rate of motion as the liquid particles because their temperatures are the same. The particles in the gas phase, however, are more disordered and move through greater distances before colliding with the walls of the flask or with the liquid surface or with one another. Some of the gas particles encounter the liquid surface in a way that causes them to travel into the liquid. They are trapped and become a part of the liquid. Such particles have condensed. In doing so they share the speed of motion and freedom of motion that they had as vapor with the liquid particles. Although the average speed of the particles in the liquid phase may have been the same as that of the particles in the vapor phase, these condensing particles warm the liquid of which they are now a part. In sacrificing the freedom of motion, the particles had excess energy to share with particles in the liquid.

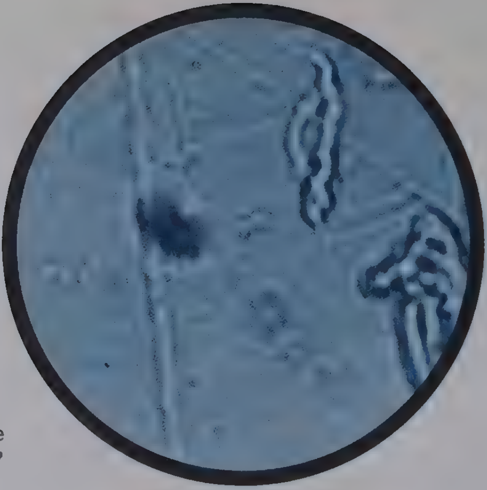
In the liquid phase particles are close together and must consequently bounce against one another much more often than the gas particles. Occasionally, however, a particle collides so as to pick up enough motion so that it can travel out of the liquid and into the gas-filled region. The particle or particles with which it collided imparted some motion (heat) to the escaping particle and are cooler as a consequence. This occurrence must be a special kind of collision since the collision does not result in an equal distribution of the motion of the colliding particles. The escaping particle may gain motion because the collision is with two or more particles simultaneously; the motion gained may be enough more than the average so the particle can escape from the surface.

The breathtaking experience in the skating game called "crack the whip" illustrates this transfer of motion. Here, a line of skaters with hands joined speeds along with the individuals nearly abreast. The pivot skater on one end of the line turns sharply and stops quickly, pulling on the next in line. This draws the next skater into a slightly larger turn and he will find his speed a little greater than before. Some of the forward motion of the pivot skater was transferred in this act of changing direction. Each skater along the line, if the members do not weaken and lose their grip, gains some of this speed and a little more from each one between him and the pivot. Finally, if there are six or eight in the line, the skater at the tip of the "whip" may be unable to hold on for his speed is too great to turn so sharply. When this occurs, he will "evaporate" from the more orderly group. The motion of the members of the group is less (they have "cooled") in order to evaporate the one.

At a particular temperature the average amount of motion of the particles determines how many will escape into the gas phase. At a low temperature, a smaller fraction of the total number may gain the motion necessary to escape. The slower-moving particles cannot contribute as much spare motion to the one that eventually is freed. The probability of the proper conditions existing is much less in the colder situation. There is, according to this probability, a constant rate of escape for any given temperature, but the rate of escape is so great at the point where the heat is applied that bubbles of vapor appear within the liquid. $\boxed{10}$ Did you notice any change in the size of the vapor bubbles during boiling in the sealed flask? $\boxed{11}$ Would a change in size be important to our effort to develop a model?

If the boiling liquid is in a closed flask, the number of gas particles increases. As the number increases, the prospect of some of them going back into the liquid also increases. When the rate of return to the liquid (condensation) equals the rate of escape (evaporation), equilibrium exists. [B, C, D]





Does evidence for particle motion exist in the microphoto of the milk drop?

7–4 THE TWO HEAT THEORIES AND DIFFUSION

7 3 If we can gather some evidence that the particles of a liquid are in fact in motion and that the motion is related to the temperature, it would make the particle theory more acceptable.

For a simple experiment you will need a microscope, a slide, a cover slip, and a few drops of very diluted milk—so dilute that it will allow a considerable amount of light to pass through it. Place a drop of the diluted milk on a microscope slide and lower a cover slip onto the drop to make a thin film of this suspension. Examine the drop under your microscope with high magnification and bright lighting from below. Determine whether the suspended milk particles show the same erratic vibrations as we observed for similar particles suspended in gases. Warm the slide and see if this affects any observable vibration.

Watch the diffusion of a few drops of milk through a glass of still water. 12 Does the diffusion rate seem to be influenced by temperature? 13 To what extent is the mixing process affected by convection currents?

7 4 For another simple experiment you will need two tall containers, hot and cold water, and a few crystals of potassium permanganate. Drop one tiny crystal of potassium permanganate into a tall container of cold water and one into a tall container of hot water, and observe what happens. 14 Does this activity tend to support one heat theory more than the other? 15 If we assume that the diffusion rate depends upon rate of particle motion, can we be sure this favors either of our proposed theories?

What factors affect the rate of diffusion?

The change in rate of motion of particles seems to be an adequate explanation for differences in temperature. If so, then the particles in solids must be in motion. [16] If such motion does occur, should not a solid substance show a gradual change in shape over long periods of time? Recall the change in an alum crystal that occurred when it was placed in a solution of chrome alum in Experiment 4–4. [17] Did the color appear to penetrate the entire crystal, or was it just on the surface?

Is it possible for diffusion to take place in solids? There are no simple activities that will give us the answer, but there are some practices in metallurgy that would be difficult to explain without assuming diffusion in solids. You can read about the processes by which carbon is added to iron to form case-hardened steel, and how the semiconductor devices of the electronics industry are made by diffusing low concentrations of impurities into silicon or germanium.

7–5 THE TWO HEAT THEORIES AND SOLUTION AND CRYSTALLIZATION

We have learned that certain solids cool a liquid as they dissolve in it. [18] Is this consistent with our guess that particles of matter in the fluid state have increased their freedom of motion?

While we are thinking about particle motion and changes in temperature, we should reconsider some of the observations made during our study of crystals. The general situation we observed was that heat was involved in the change from the orderly crystalline pattern and that it was also involved in the formation of crystals. In some solutions heat was released; in others it was absorbed. With ice, heat was added which did not change the temperature but simply changed the phase. During solidification, either of hypo or of ice, the heat released was available to warm the substance from its supercooled condition to the freezing temperature. Whatever is responsible for the binding of one particle to another in crystals is related to heat. If we find that our caloric fluid theory is not adequate to explain all that we may observe, then we must find some relation between the breaking of the binding forces between particles of matter and heat. [E, F]

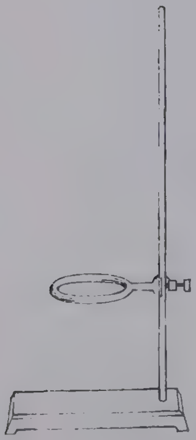
7–6 THE TWO HEAT THEORIES AND CONDUCTION

Except for the selection of certain materials as the most suitable for our calorimeters, we have not studied the transfer of heat very thoroughly. We are most concerned with conduction, the process by which heat seems to travel within a substance or between two objects that are touching each other. Recall that you found that the best insulating materials were very porous while those that were nonporous were the best conductors. If heat is a fluid, it does not travel well through the larger holes in matter but is associated closely with the particles, so that the denser substances, with their particles more tightly packed, are best suited for transferring the motion through collisions.

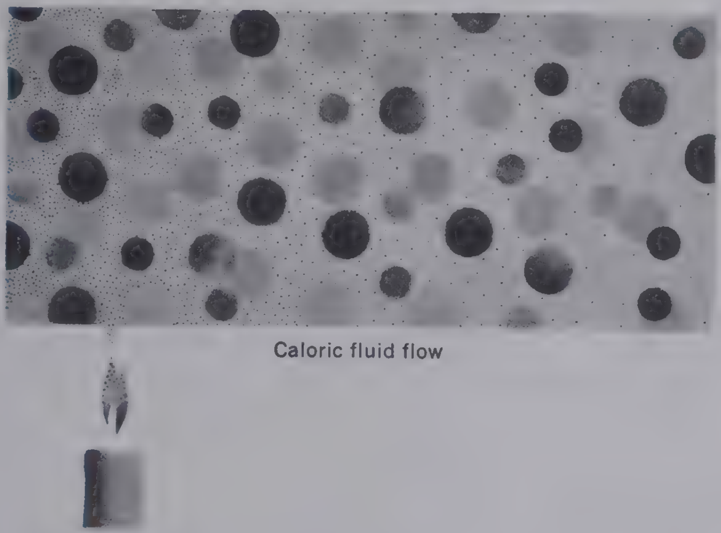
One conspicuous problem with a theory of particle motion, in which the temperature is directly related to the rate of motion, would seem to be that a substance that does not lose heat to its surroundings stays at a constant temperature. Despite the countless collisions that must be taking place, the average rate of motion does not change. This is contrary to motions we have observed of larger objects. Apparently, the submicroscopic bits of matter we believe exist may collide and rebound with no loss in total motion. Later, when we consider the topic of motion in Chapter 9, we may be better able to understand this unique situation.

TAKING INVENTORY

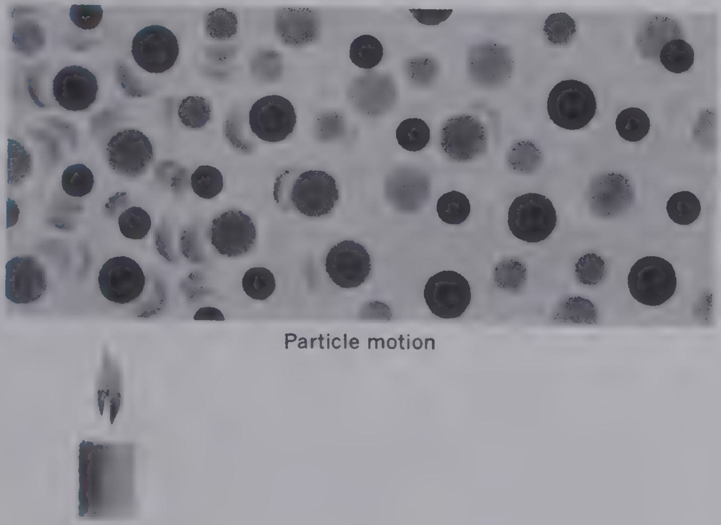
We have summarized our observations on heat from two points of view—the caloric fluid theory and the particle motion theory. We have used both theories to explain evaporation, boiling, diffusion, solution, crystallization, and conduction. In each of these processes heat has been transferred directly. It has been taken from flames or the stream of gases coming from a flame. Neither of our



Do both theories provide adequate explanation for the conduction of heat in solids?



Caloric fluid flow



Particle motion

How does heat from the sun get to the earth's surface?

theories explains the transfer of heat from a fireplace, an incandescent lamp or the sun. Such heat travels outward from the flames of a fireplace, not upward with the hot gases. It passes through the glass bulb of the lamp, even though we know glass is not a good conductor of heat. It comes to us through empty space from the sun. Surely neither a caloric fluid nor a particle motion theory is of itself sufficient explanation for these phenomena. In the next chapter we will try to determine the rules that govern this method of heat transfer. We will see how we must adjust our theories to accommodate this form of transfer.

FURTHER EXPLORATIONS

A. In order to reduce the evaporation from lakes and reservoirs in arid regions, it has been found that small quantities of certain materials will spread to a minute thickness over the water surface. These are referred to as "monolayers" because they are only one (mono) particle in thickness. While this practice does retard evaporation, it permits the exchange of gases from the atmosphere to the water and vice versa so that fish and other animals and plants may live in the water. Read about these layers and try to relate their properties to the theories for matter and energy we have been studying. Will a film of mineral oil retard the evaporation of water? We know it will arrest the transfer of gases, because applying enough of this oil to a pond will kill all living things in the water.

B. Releasing certain gases from high pressures can result in extremely low temperatures. Carbon dioxide and ethyl chloride become gases at such low temperatures that the release of these gases may produce freezing temperatures and may even damage living tissue. Ethyl chloride is sometimes used by doctors to "freeze" tissue to anesthetize it for minor operations. Such gases must absorb heat from the objects they cool. How can these facts be explained on the basis of your theories of heat and the structure of matter?

C. Refer to the Further Explorations in Chapter 5 which deal with boiling. Can you suggest useful ways of dehydrating (removing the water from) solutions or suspensions without intensive heating?

D. If fresh meat or vegetables are dehydrated they may be stored for indefinite periods without spoiling. When the water is restored they again seem almost the same as when they were fresh. Suggest other ways of drying foods without altering the characteristics any more than necessary.

E. Is there any evidence that a caloric fluid might be responsible for the changes that take place in food when it is cooked (vegetables becoming softer, eggs becoming harder, colors often changing)? What other explanations can you advance for these changes?

F. Is it reasonable to suggest that bacteria and fungi live on caloric fluid? Hint: Freezing food retards its rate of spoiling or molding.

FURTHER READING

BORN, MAX, *The Restless Universe.* Dover Publications, New York, 1951.
Comprehensive treatment of molecular motion, kinetic theory, heat conduction, and other topics related to Chapter 7.

BRAGG, WILLIAM, *Concerning the Nature of Things.* Dover Publications, New York, 1948.
Contains pertinent information relating to heat.

JONES, HOWARD M., and I. BERNARD COHEN, *A Treasury of Scientific Prose: A Nineteenth-Century Anthology.* Little, Brown and Company, Boston, 1963.
Another source of insight into the thoughts of scientists, available through their own writings.

Lives in Science, by the editors of Scientific American, Simon and Schuster, Inc., New York, 1957.
A source of information on how pioneering scientists studied the problems similar to those in this course.

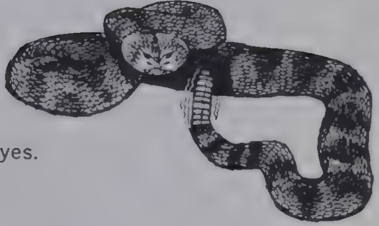
PATTON, A. R., *Science for the Non-Scientist.* Burgess Publishing Company, Minneapolis, 1962.
Gives a good background in what science is and would be good reading at this point in the course.

ROLLER, DUANE, editor, *The Early Development of the Concept of Temperature and Heat: The Rise and Decline of the Caloric Theory.* Harvard University Press, Cambridge, Mass., 1950.

RUCHLIS, HY, *Discovering Scientific Method*, Harper & Row, Publishers, New York, 1963.
Contains examples of how scientists solve problems. Gives the student a chance to compare personal experiences in science with those of scientists.

8 Radiation

8–1 Detecting Radiation
8–2 Relationships Between Temperature and Radiation
8–3 Separating Light From Heat Radiation
8–4 Radiation and Plants
8–5 What Is Radiation?



Notice the pits of the rattlesnake just below the eyes.

8–1 DETECTING RADIATION

The warmth you feel from a fireplace or the sun comes to you by a means called *radiation*. The caloric fluid theory is severely strained as an explanation for radiation when we realize that rays from the sun or a fire in a fireplace pass through glass: We have found glass to be a poor heat conductor. On the other hand, a moving particle theory does not explain how radiation can reach us from the sun if there are no particles in space.

Neither of these theories is able to account for the fact that radiation apparently travels from the source over any distance and instantaneously. Radiation takes place best in a vacuum or through transparent materials. In fact, the word *transparent* means able to transmit radiation. The opposite of transparent is *opaque*. Opaque objects do not allow radiation to pass through them.

Rattlesnakes, cottonmouth moccasins, and copperheads have small indentations, or pits, in front of each eye that seem to be very sensitive to differences in radiation. These snakes can detect radiations from a warmer object, even if it is only a few degrees warmer than its surroundings. Thus, even though a rattlesnake's eyes become cloudy when it is shedding its skin, it can detect the presence of warm-blooded animals. This gives the snake protection against enemies and enables it to find food like mice and squirrels.

Nerves in our skin respond like the pits of snakes but with far less sensitivity. You have sensed heat radiation from a light bulb. You may have noticed the warmth of a wall covering a steam pipe as you pass it. In addition to feeling radiation coming to our bodies from a warmer source, we are also aware of heat going out from our bodies to cold surroundings. We ordinarily draw draperies across a window on a cold night. They not only provide privacy (by being opaque to the light that would leave the house) but also reduce the radiation from the warmer objects inside to the cooler region outside. This provides evidence that all objects are constantly engaged in an exchange of radiation with all other objects. Even when two objects have reached the same temperature, this exchange continues, in still another example of equilibrium.

$\boxed{1}$ If you do not know whether an electric iron is plugged in, how close do you have to get to it before you can tell whether it is hot? $\boxed{2}$ Are there certain areas of your skin that are more sensitive than others to radiation? In the days before automatic stoves, an experienced baker could tell whether the oven was ready for baking by putting his hand inside it. $\boxed{3}$ Was the baker measuring air temperature or the radiation from the oven walls? $\boxed{4}$ Does the temperature of the air or the radiation cause pastry to brown? Consider a toaster in arriving at a possible answer. $\boxed{5}$ Does the amount of moisture in the surface layer of a slice of bread make a difference in the browning? To be scientific, you need to design experiments which can lead to clear answers to these questions.

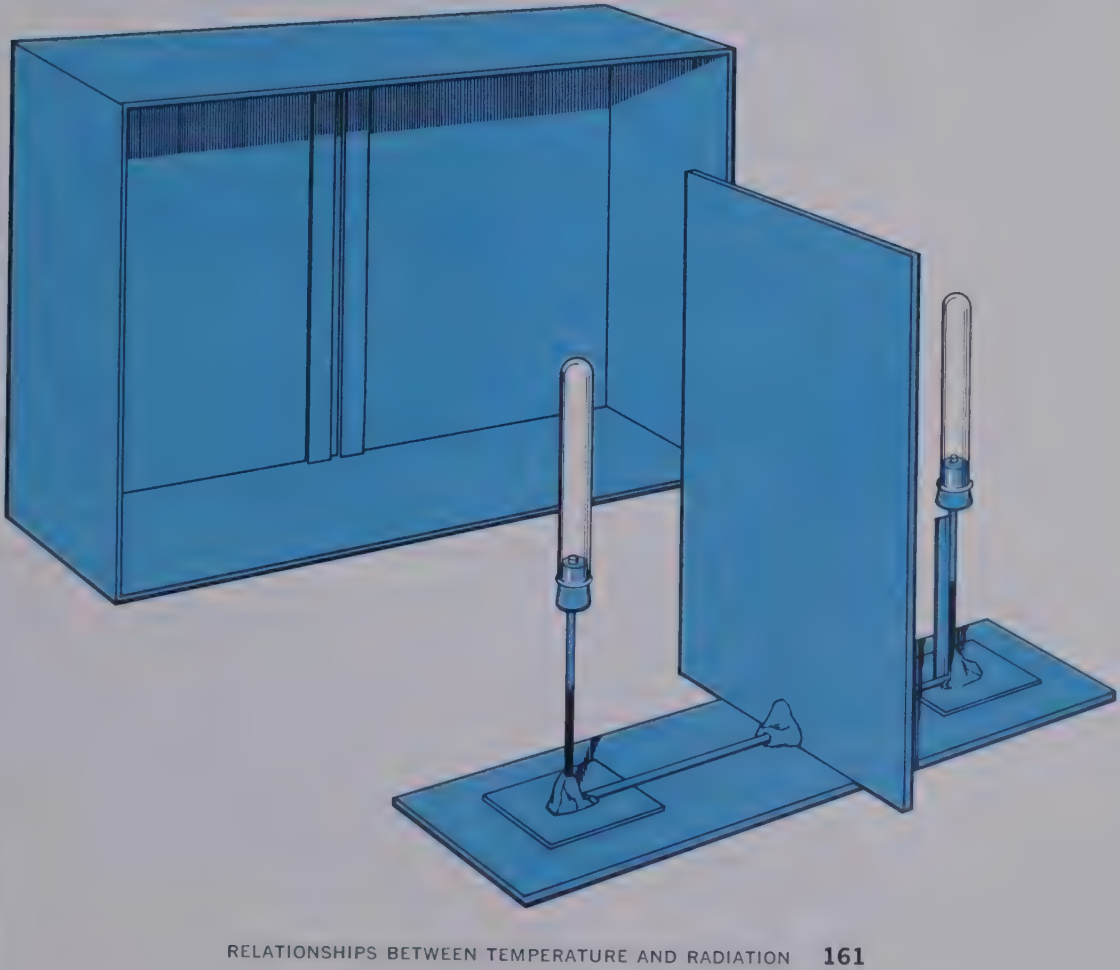
8 1 Devise a blindfold from several folds of black cloth which exclude all light. Your teacher or a fellow student will lead you near incandescent and fluorescent bulbs, a hot iron, and direct sunlight. Try to identify each source of radiation. Still blindfolded, see if you can detect a can of ice cubes, a can of room-temperature water, and a can of hot water when each can is only a few centimeters away. $\boxed{6}$ Why should you use a black cloth for your blindfold? $\boxed{7}$ Is an object that appears black to our eyes also opaque to other radiations as well as to light? [A]

8–2 RELATIONSHIPS BETWEEN TEMPERATURE AND RADIATION

Does the rate of radiation depend on temperature? Again, our nervous systems are not adequate for quantitative study, but we can devise an instrument that is sensitive enough to detect small differences in temperature. This device is called a *differential air thermometer*.

8 2 To make a differential air thermometer you will need a length of U-shaped glass tubing, two test tubes, water, a heat shield (such as a large piece of corrugated cardboard), and two one-hole stoppers to fit the test tubes. Pour water into the U-shaped tube. Place a one-hole stopper in each end of the U-shaped tube, and put a test tube over each stopper. Since this is a sealed system, it is not influenced by outside air pressure and the relative level of the liquid in the two legs of the tube will be determined by the temperature of the gas in each test tube. Place the heat

The differential air thermometer.



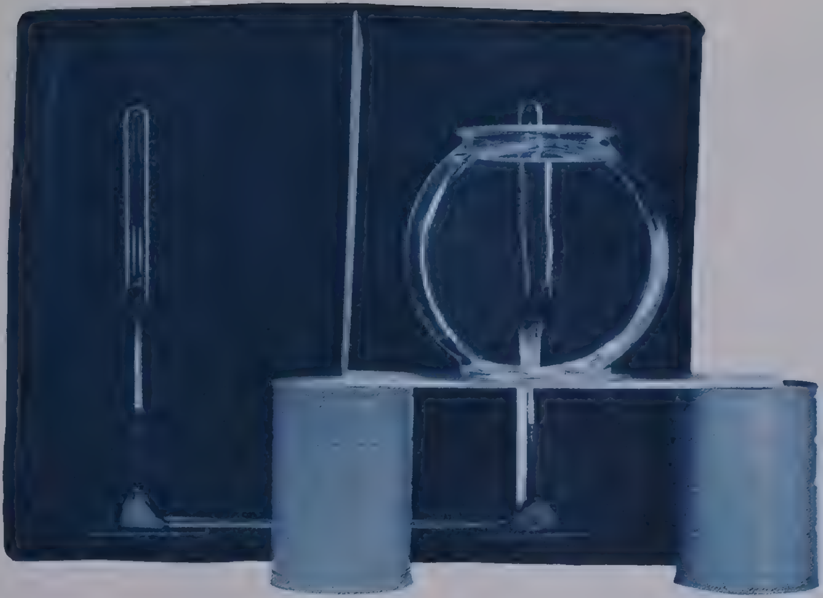
shield between the test tubes so that they are isolated. Surround the system with another shield, such as a cardboard box open at both ends, to isolate it from other heat sources.

With this apparatus, we can make use of our knowledge that the volume of a gas is related to the temperature. In Experiment 6–2, the heated gas expanded its volume in an unrestricted way. The atmosphere exerted the same pressure after heating as it had before heating. Thus the temperature change was apparent as a volume change. In the differential air thermometer, the temperature change will show as a volume change but it will not be as great, since the gas is not expanding against a constant pressure. As the gas expands in one side of the thermometer it is compressed in the other side. As a result the pressure rises in both bulbs, on one side because of increased particle motion and on the other because of decreased volume. The interaction of these three factors—temperature, pressure, and volume—are such that we cannot readily calibrate this device to show difference in temperature by degrees, but we can show generally that temperature varies with radiation rates.

Place a metal container of hot water near one of the test tubes and observe the column of water. [8] Why would the calorimeters we used in Chapter 6 be unsuitable for this experiment?

To establish clearly that any change in the column was due to heat gained from the hot container, place two similar containers, one near each bulb, and observe the indicator position. Finally, fill one of these containers with hot water and place it a few centimeters from one bulb; fill the other container with cold water and place it at the same distance from the other bulb and again observe the indicator position. Record, at regular intervals, the temperature of each of the two containers of water as they come to room temperature. Also record the change in position of the indicator column of the differential air thermometer. Make a graph of the position of the indicator column as it is related to the difference in temperature.

Repeat the experiment, using hot water in *each* container again, but vary the distance between one of them and the bulb of the thermometer to try to determine any relationship between the distance and the transfer rate. What information do you need to record? The temperature of the water should be kept as high as possible, but it should be the same in both containers. [9] What is the effect of changing the area of the radiating surface? Consider the influence of the radiating surface by using one bright container and one dark container for the water. Be sure they are about the same size and at the same distance from its bulb. Explore the influence of the color of the thermometer bulb by covering one test tube with bright aluminum foil or soot from a candle.



Is the water opaque to radiant heat rays?

We know that radiation is not a simple thing to study. When it comes from a fire or incandescent bulb, it seems to be both heat and light. Radiation from a fluorescent bulb seems to be much more light than heat. Radiation from an electric iron seems to be all heat, since we cannot tell by looking at it whether it is hot or not. [B, C, D, E, F]

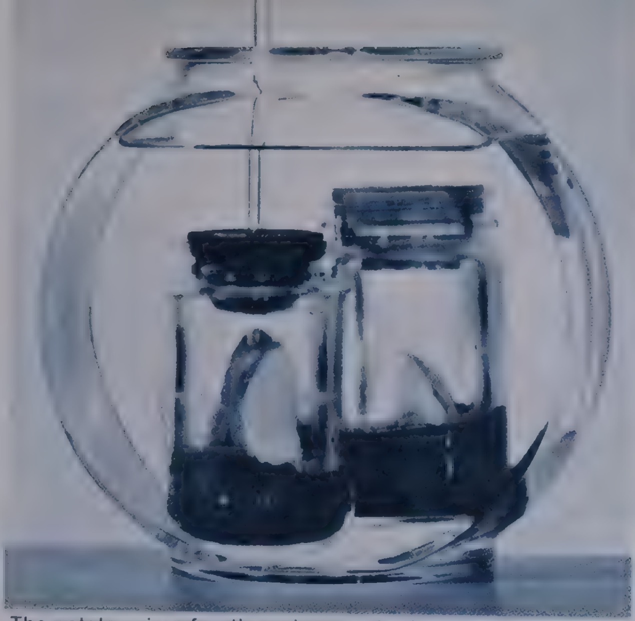
8–3 SEPARATING LIGHT FROM HEAT RADIATION

There seems to be a difference between light—the radiation that enables us to see—and the radiations that cause a change in temperature. In general usage, the word *light* refers to that radiation to which our eyes are sensitive. Do some materials separate heat from light radiation? Place a narrow aquarium or other straight-walled container of water between a glowing light bulb and one of the tubes of the differential air thermometer. [10] Do you detect a difference in the amount of heat reaching the bulb and the tube? [11] What happens to this heat? [12] Is there a similar way we could check materials to determine if they allow heat to pass through and not a great amount of light? [13] Is heat, like light, reflected by a polished metal surface? [14] If a cylindrical water container is used instead of the rectangular one, what difference, if any, would you expect?

8–4 RADIATION AND PLANTS

Plants can help us in our study of the nature of radiation. For an experiment with sprouting potatoes you will need a potato, some soil, a glass jar, three beakers or plastic cups, and two large cans, one with a plastic top. Cut the potato into four pieces. Save the fourth piece for later. Plant each piece in soil in a cup and water it. Invert the glass jar over one cup and one of the cans over another. Put the third cup in the other can and cover it with its plastic cover. Put all three setups in an area where they will receive light, preferably sunlight, at relatively regular intervals. From time to time examine the shoots. Notice the length, direction, color, and other characteristics of these young plants. [15] Can you draw any conclusions about the effects of light and of radiant heat on the growth of plants? Try similar experiments using tubers, bulbs, or seeds.

We know from having observed the opening and closing of flowers that plants seem to respond to the stimulus of radiation. The direction in which the leaves and blossoms of plants are pointed changes as the sun's position in the sky changes. This is particularly noticeable with



The potatoes in a fourth environment—the aquarium. What is the variable between the two jars?

An algae aquarium.

beans and sunflowers. Are the plants responding to the visible light from the sun or would they similarly respond to a moving source of heat alone? Add a fourth part to the experiment with the potatoes in the form of an environment where the pieces are in clear glass jars submerged in the water in an aquarium or fishbowl. Most of the radiant heat cannot reach such a jar. [16] What influence does this environment have on the growth of the shoots? [17] Is there a relationship between the growth and the temperature when all other factors are kept the same?

8 5 Perhaps you have noticed the green algae that may form on the sides of an aquarium receiving direct sunlight. Screen off the direct light from part of one side of such an aquarium with opaque paint or paper and try to determine if this growth depends upon the heat or upon the light that acts on the water. Attach horizontal strips of cellophane of different colors to the outside of the aquarium to determine if the growth of this tiny organism depends upon the color of the light. Scrape specimens of algae from the sides of the aquarium; obtain a specimen from each kind of environment. Examine the specimens with a magnifier or a microscope to determine if there are differences in the growth pattern in the different light regions.

8–5 WHAT IS RADIATION?

Throughout this chapter we have been observing the effects of radiation on living and nonliving matter. Radiation has shown itself only by these effects; we have no direct evidence of its presence. We could continue to study radiation by observing its effects on light-sensitive materials such as photographic film or paper. In these materials changes take place which depend on the kind and amount of radiation that strikes the material. The changes are so subtle that they must be detected by a process called *developing*. In the heat-sensitive nerve endings of our skin or the retinas of our eyes, slight changes take place as long as radiation is received.

Does radiation consist of matter in the sense that we have used the word matter? We cannot measure it in the usual way or describe it as either a solid, liquid, or gas. Radiation is, in fact, not matter since it does not have mass or take up space. It calls for a different classification. Radiation is a form of energy. The word *energy* is derived from a Greek word meaning work. Energy produces changes, and changes do not result unless some kind of energy is spent—unless work is done.

8 6 We can see radiation doing mechanical work (that is, moving matter) by using a device called a radiometer. Place the radiometer in bright light. 18 Would the radiometer operate as well, or at all, if it were not protected by the glass case? 19 Does its rate of rotation vary with the amount of radiation? Use incandescent lamps of various wattages at a constant distance and count the rotation rate of the vane for each wattage. 20 Is the differential thermometer as sensitive to the different wattages as the radiometer? 21 What happens to the rotation rate if you add a reflector to a bulb or use a bulb of comparable wattage with a built-in reflector? 22 Will a heated electric iron produce the same result as a lamp of comparable power? Wattage is a measure of the rate at which energy is being used in the device. This is another way of saying the rate at which some kind of work is being done. The rotation rate of the vane in the radiometer is thus a measure of the work being done; if it changes from one test to another, this is an indication that the radiometer is receiving different amounts of radiant energy.

Now that we have studied heat transfer by contact between hot and cold objects (conduction), by streams of gases or liquids (convection), or through space and certain materials in which neither conduction nor convection applies (radiation), we can refine our definition of heat. To the specialist, the word "heat" is used only when energy is actually being exchanged. Technically, then, we should not speak of the heat energy "in" an object. We should only speak of heat energy as being transferred from one object to another. [G, H]

The radiometer is an example of conversion of radiant energy to mechanical energy.

TAKING INVENTORY

Radiation produces changes in the objects it strikes. We consider it to be a form of energy, for only when some energy is exchanged can change of any kind take place. Some substances transmit radiation, others reflect it, and many absorb it.

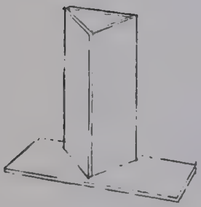
The nature of radiation cannot be explained fully by either the caloric fluid theory or the particle motion theory for heat. Radiation, however, is too important a part of the world around us to be ignored, even if it is not readily explainable. We have gained experience with the interaction of heat radiation and motion through our studies in this chapter. In Chapter 9, we will study the energy of motion in more detail. As you study it, watch for ways in which motion and heat are related. Can motion be translated into heat or heat into motion?

FURTHER EXPLORATIONS

A. Derive a definition of the word *black* based on our study of radiation.

B. Attempt to redesign the differential thermometer so that you can use it to detect radiation from horizontal surfaces. With such a redesigned instrument you can examine the radiation and heat loss from the liquid surface in a calorimeter as well as from its sides.

C. A prism will disperse white light into the spectrum of color. Place a large prism in sunlight and hold the bulb of a thermometer in various bands of color for a time to be certain it is showing any temperature change resulting from radiation of that color. Are there any regions beyond the visible spectrum where the temperature also seems to be affected? Try using the differential air thermometer as the radiation detector. Try blackening the bulbs of both thermometers and repeat this experiment.



D. Water prisms can be built from microscope slides and waterproof cement. Four slides are necessary, three for the upright sides of the prisms and one for a base. The prism can be held together with tape while the cement dries. Build a water prism and use it instead of a glass prism in the experiment in Further Exploration C.

E. Obtain a concave reflector, 20 cm or more in diameter, such as one finds in an automobile headlight, and use it to concentrate the light from a candle or lamp onto the bulb of a thermometer. Determine the characteristics such a reflector must have to provide the greatest concentration of heat from a certain source at some particular distance. Consider the diameter, curvature, reflecting surface, and any other factors that seem to make a difference.

F. We have said that calorimeters are poor radiators of heat. Test the validity of this statement by attempting to detect radiation from some of the containers you considered for use as calorimeters.

G. How would you determine the temperature of a sheet of transparent material (glass or plastic) heated by sunlight? Does glass absorb any light? How could you determine the amount of radiant energy a unit area of glass absorbs? Could you determine the amount of reflection from a flat glass surface? Is this reflection the kind of radiation that activates the radiometer?

H. You may have used a magnifier to concentrate the radiation from the sun on paper or wood to cause a fire. You could also use a magnifier to focus the radiation from a candle flame or a lamp on the bulb of a thermometer. You will be able to see where the image of the visible flame is brightest, but is this the spot where the greatest amount of radiant heat is concentrated? Could you use a radiometer to test for the greatest amount?

FURTHER READING

ADLER, IRVING, *The Secret of Light.* International Publishers Company, New York, 1952.

An explanation of basic principles of light for the student who wants to explore the topic further.

RUCHLIS, HY, *The Wonder of Light.* Harper & Row, Publishers, Inc., New York, 1960.

A thorough treatment of light. Contains many photographs and intriguing comparisons to hold the reader's interest.

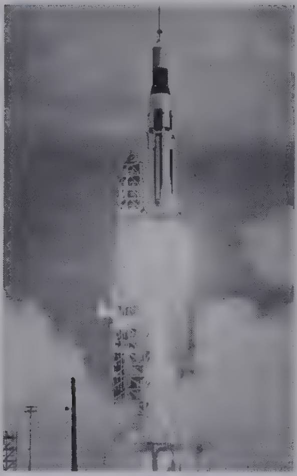
TANNENBAUM, BEULAH and MYRA STILLMAN, *Understanding Light: The Science of Visible and Invisible Rays.* McGraw-Hill Book Company, New York, 1960.

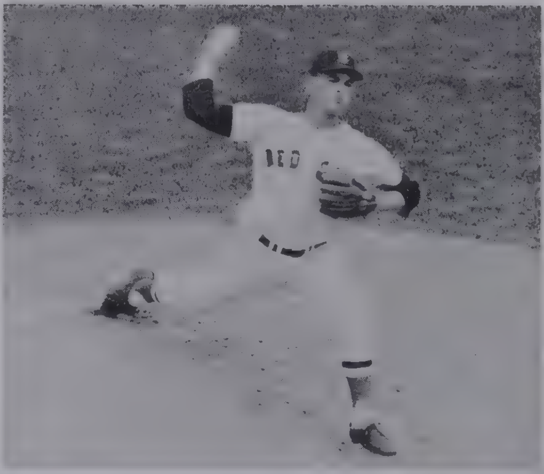
A very comprehensive treatment of light. Relates many modern topics to the work of pioneer investigators in this field.

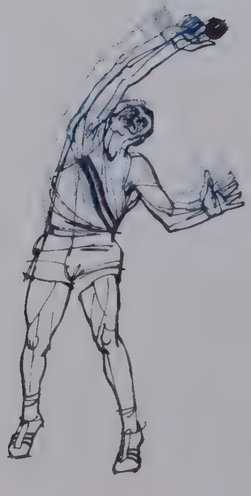
WILSON, MITCHELL and the Editors of Life, *Energy.* Time-Life Books, Time, Inc., New York, 1963.

9

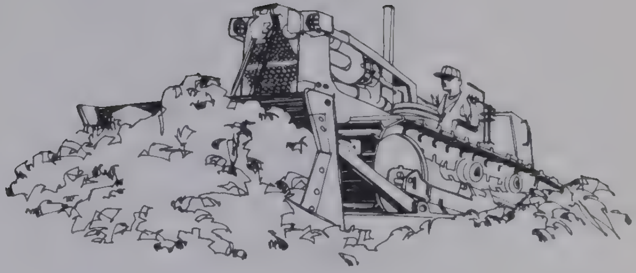
The Energy of Motion







9–1 Force, Distance, and Energy
9–2 Motion Energy and Position Energy
9–3 Levers and the Conservation of Energy
9–4 Equilibrium of Forces in Levers
9–5 Levers as Simple Machines
9–6 The Earth's Gravitational Force
9–7 Energy Stored in Elastic Materials
9–8 Partially Elastic Materials
9–9 Elastic Properties of Liquids and Gases
9–10 Is Energy of Motion Always Conserved?

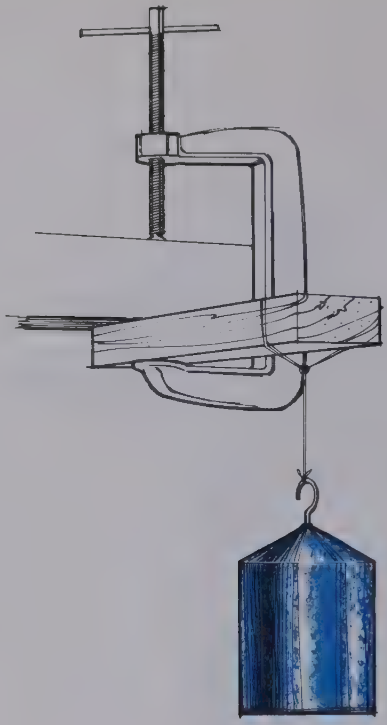


9–1 FORCE, DISTANCE, AND ENERGY

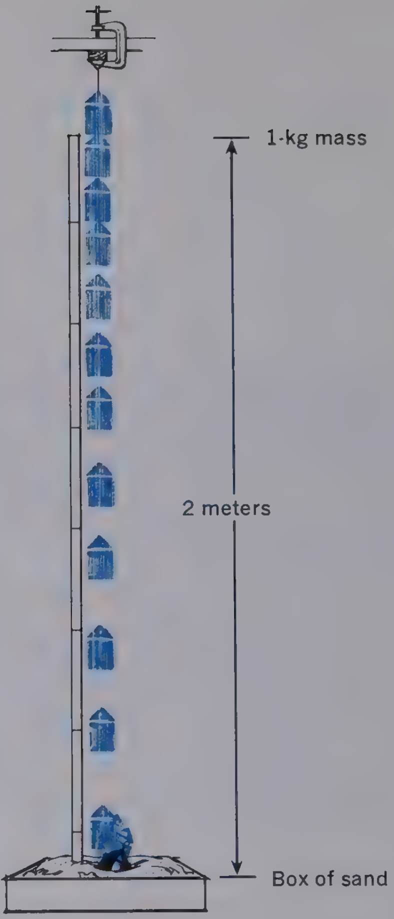
The delicate, rotating vane of the radiometer and a bulldozer used in road construction seem vastly different. But both are machines that convert energy from one form to another. Any device that modifies, transforms, or transmits energy is a *machine*. A machine may be simple or complex. The radiometer converts the energy carried by the beam of radiation into motion of the vane. The chemical energy of the fuel is converted in the bulldozer into energy of motion of the machine and its burden. This energy of motion is called *kinetic energy* from a Greek word meaning "motion." It is one of the most evident forms of energy because we can see the motion or feel or see the impact of a moving object and the changes that it produces.

To give an object kinetic energy—to make it move—a force must be applied. From experience you probably realize that the amount of energy given an object depends, in part at least, on the amount of force used. Another factor, which you may not have noticed, is that the amount of energy also depends on the distance through which the force acts.

For our discussion of the energy of motion we will use as an example an object having a mass of one kilogram (kg). You will recall from your work in Chapter 2 that gravity exerts one kilogram-force (kgf) on this object. If you release this object, it will fall in response to this force of gravity. It will gain speed until it encounters an arresting force; then it will stop. The object gains in kinetic energy as its speed increases. The amount of kinetic energy that the kg mass can acquire depends on the distance through which the force of gravity is able to act.



Gravity is the mutual attraction of the 1-kg mass and the earth and produces 1 kgf.

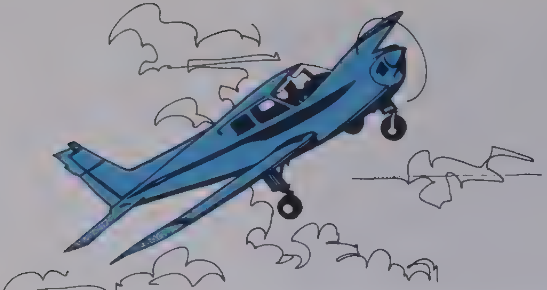


If an object of 1.0-kg mass is released from a point just two meters (2.0 m) above the floor, at the moment of striking the floor it will have kinetic energy of 1.0 kgf × 2.0 m, or 2.0 *kilogram-force-meters.* The word *work*, as used in science, is defined in the same way as energy and measured in the same units, that is, in kilogram-force-meters (kgf-m). In the example of the falling body, the work input is equal to the resulting energy.

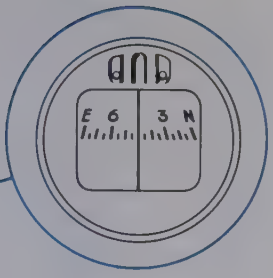
With these definitions in mind, let's study and measure the forces and motions in some simple energy transfer situations. In the discussion we will refer to the same object suggested earlier, a 1.0-kg mass. In your experimenting you may use any object that you can conveniently weigh and move. Our rules would be of little value if they applied only to special situations such as a mass of 1.0-kg or a distance of 2.0 m.

How much work is done if a 1-kg mass is lifted 2 m?

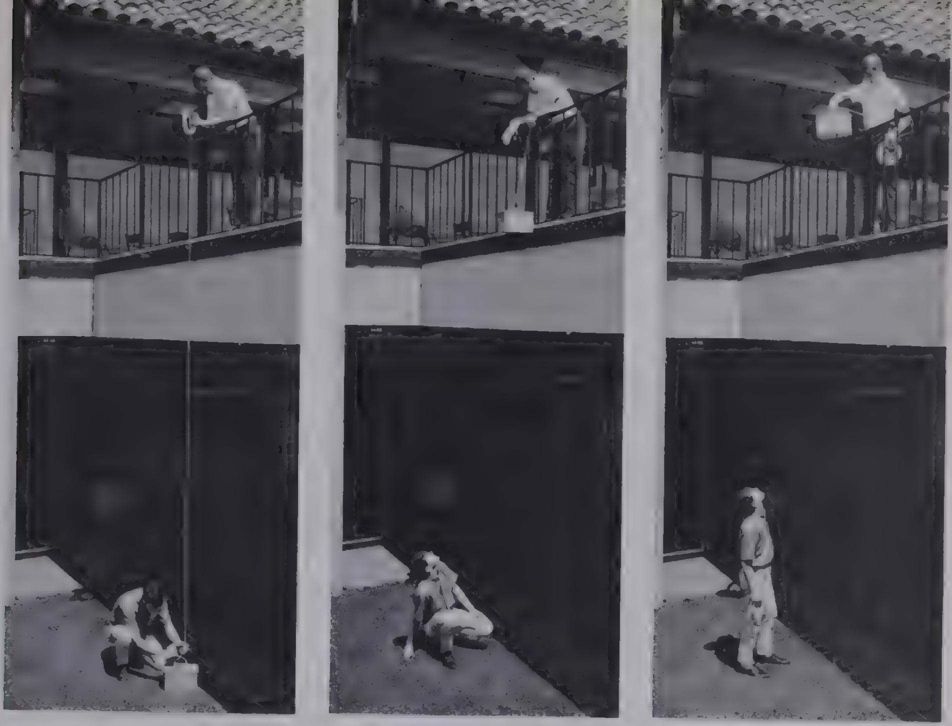
For a simple experiment with work, you will need a 1.0-kg mass and a spring balance. Suspend the 1.0-kg object by the spring balance and note and record the scale reading as you lift the object slowly and steadily, as you hold it stationary, and as you lower it slowly and steadily. [1] In terms of our definition for work, how much work have you done if you lift this object from the floor to a height of 2.0 m? Perhaps you can understand why we have chosen this definition of force times distance for both work and energy. If, in the example we used, we must pull with 1 kgf through a distance of 2.0 m to raise the object, this should also be the amount of energy of motion the object has when it has fallen freely (without any restraining force) to its original level. This means that the energy we spent in raising the object was later used to set it in motion, and no energy was lost.



Navigating an airplane requires knowledge of speed and direction.



Again use the spring balance and lift the 1.0-kg mass, straight up to a height of 2.0 m, and thus do 2.0 kgf-m of work upon it. Now from this elevated position, move the mass horizontally 2.0 m, always keeping it the same height—2.0 m—above the floor. What did the balance read as the object moved horizontally? With this force acting, you have moved it an additional 2.0 m. What would be its energy of motion at floor level if you were to release it from this second position? There is still the same energy of motion when dropped from this point, for gravity's force has acted through the same total distance, 2.0 m, in speeding it toward earth. The horizontal motion did not change the total energy due to gravity and thus it does not represent work recoverable through the action of gravitational force. During the horizontal motion, although a 1.0 kgf was being applied to the object all of the time, there was no motion in the *direction* of this force. The gravitational force was vertical and this new motion was horizontal. It is the motion *in the direction of the force* that contributes to work done. Both motion and force, then, are more than simple numbers. To describe them fully, the characteristic of direction must be specified. Quantities such as force, motion, and distance, which have both a magnitude and a direction, are called *vector* quantities in contrast to *scalar* quantities which have a magnitude only. [A]



What potential energy does the box have in relation to the first floor and the second floor in the three photos?

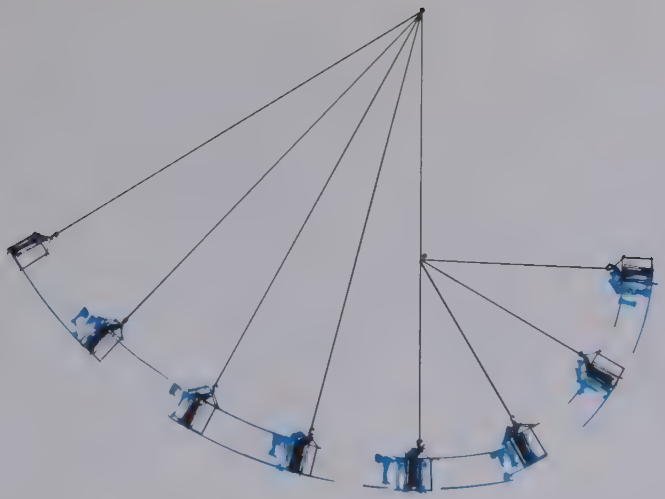
9–2 MOTION ENERGY AND POSITION ENERGY

We need to introduce another term into our study of motion energy if we are to say that energy is *always* conserved. This term will be used to describe the condition when an object is lifted to its highest position and is about to be released to fall. At this point, because of the work done to raise it, the object is capable of doing this same amount of work in falling. The word *potential* describes this capability. The object we are considering has 2.0 kgf-m of *potential energy* with respect to the floor when it is raised to this 2.0-m height. Notice that, if we are to apply numbers to a description of potential energy, we must indicate a level or reference point as having zero potential energy. In this example the zero energy level might be either the floor or some other reference level. If the experiment with the 1.0-kg mass were done on the second floor of the school, the potential energy with respect to that floor is less than the potential energy with respect to the ground outside the building. More work would have to be done to bring the object from ground level to the point off the second floor, than from the floor level of the second story.

Trace the energy condition of the pendulum as it swings back and forth.

A pendulum serves as an excellent example for our study of the interchange of position energy and motion energy. As it swings back and forth from one end of its path to the other, it constantly and gradually converts gravitational potential energy to kinetic energy and back again. At the top of the swing the energy is all potential; at the center the energy is entirely kinetic. As the pendulum passes this midpoint, the energy of motion begins to change to energy of position again as the pendulum is lifted a distance against the force of gravity. That the pendulum ball (or bob, as it is frequently called) continues in this motion for some time indicates that almost all of the energy is conserved.

9 To do a simple but convincing experiment that suggests that energy is conserved as it changes from potential to kinetic and back in pendulum operation, you will need a rubber stopper, about 1 m of strong, lightweight thread, a chalkboard and chalk, and a pencil with an eraser on one end. Make a pendulum of the rubber stopper and the thread. Suspend it from the top of the chalkboard in such a way that it will swing along the board without dragging against the surface of the board. Draw a horizontal line on the board at the lowest level where this pendulum will swing, and another line 20 cm above and parallel to the first. Hold the stopper out along the board so that it is raised to the level of the upper line while still keeping the string tight. Release the stopper and notice the details of its path, particularly the height to which it rises at the other end of its swing, and again on its return to the end where it was released. When you are confident that you can predict the height to which

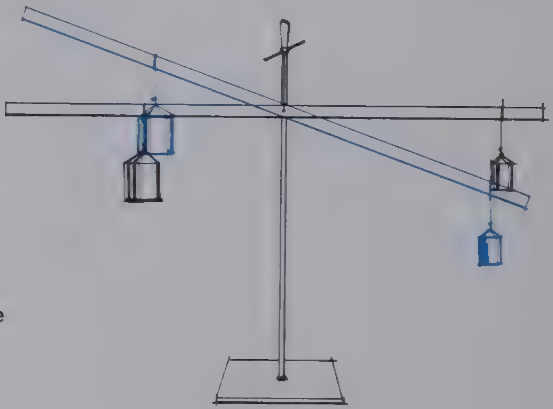


the bob will rise on any particular swing, place the eraser end of the pencil against the chalkboard so that it stands out across the arc of the string and interrupts the regular motion of the pendulum. Notice the height to which the stopper now rises as it moves with this new restriction.

Repeat this experiment many times, placing the pencil in the path of the string at various places, changing both the effective length of the string and the position in the arc where the motion is changed. With each swing, note the approximate height to which the stopper now rises and compare it with the height you would have expected it to reach if it had not been disturbed. ☐2 How do the results of this experiment illustrate conservation of energy?

As you record your observations in your laboratory record book, you should use commonly accepted words to describe motions and positions. The entire motion of a free-swinging pendulum from one side to the other and back again is called a *cycle*. The time required to complete one cycle is known as the *period* of the pendulum. Both the words *period* and *cycle*, when used in this way, carry the idea of a repeating situation, and pendulum motion is often described as cyclic or periodic motion. The motion from one extreme to the other, or from the lowest point to one end and back to the midpoint, is one-half of a cycle. The maximum distance that the bob moves out to either side from the point of lowest potential energy is referred to as the *amplitude* of the motion, and this amplitude will be seen to change slightly as the pendulum swings over a period of time. ☐3 Does the period change as the amplitude becomes successively smaller? [B]

A bicycle is a typical example of a mechanical lever system.

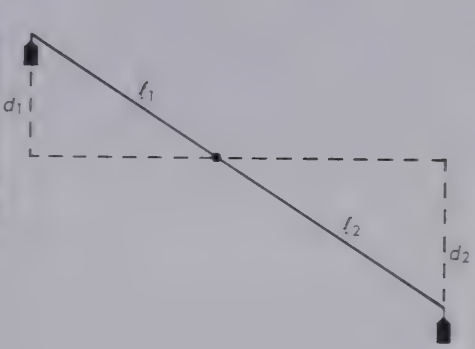


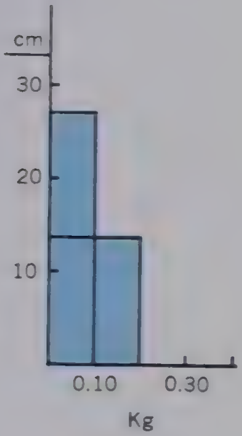
How far must a 100-g mass be placed from the fulcrum to balance a 200-g mass at 20 cm?

9–3 LEVERS AND THE CONSERVATION OF ENERGY

With a lever, we can lift objects several times our own weight. It would seem that we get more work out of a lever than we put into it. Is the operation of a lever in conflict with the principle of energy conservation we encountered earlier? A simple experiment should help us answer this question.

You will need a meter stick to serve as a lever, a piece of thread, one 100-g mass, one 200-g mass, and one 500-g mass. Support the meter stick at its balance point with a piece of thread. At a distance of 20 cm from the pivot, hang a 200-g mass and balance the system by suspending a 100-g mass on the other side with a piece of the thread. Record the





distance of this second object to the pivot, or *fulcrum*, as it is commonly called. Move one side of the balance up and record the distances one mass moves up and the other mass moves down. Work is being done by or on each of the masses during this operation. The mass that descends, responding to the force of gravity acting on it, does an amount of work equal to this force multiplied by the vertical distance it moves. Remember, work = force × distance. The rising mass has work done on it equal to the product of the force necessary to overcome the pull of gravity on it and the vertical distance it moved. The bar graph illustrated shows the work done by or on each of these objects in a typical test. In this instance, the 200-g mass moved 13.6 cm while the 100-g mass moved 27.2 cm. $\boxed{4}$ How does the work done by the falling mass compare with the work done in the opposite direction on the lifted mass?

You have in your records the measurements corresponding to d_1, d_2, ℓ_1, and ℓ_2 as shown in the illustration above. Compare the ratio d_1/ℓ_1 with the ratio d_2/ℓ_2. $\boxed{5}$ Is there a relationship between the two ratios that is characteristic of all combinations of masses that are in balance? $\boxed{6}$ Is energy conserved in this lever system?

It is evident that the lever does not help us gain energy. It simply allows us to use the energy we do have to do a task we could not do without the machine. It influences the force we can exert, but does not increase the energy beyond the quantity we invest.

Suspend the 500-g mass from the meter stick in place of the 100-g mass and move it or move the 200-g mass until the system is again balanced. Repeat the measurements made earlier as the lever is moved through one half of an up and down cycle. Calculate the work done by, or on, the 200-g mass by gravity and divide this by the distance through which the second object moves. If our conservation principle is valid, this should be a measure of the object's mass. $\boxed{7}$ What is the ratio of the longer lever arm to the shorter? $\boxed{8}$ What is the ratio of the larger mass to the smaller mass as it was calculated by the energy conservation principle?

9–4 EQUILIBRIUM OF FORCES IN LEVERS

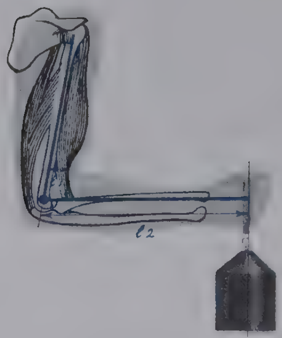
In studying the lever system we have measured the distance moved by each of the masses and found that $w_1 \times d_1 = w_2 \times d_2$. The force times the distance it moves gives us the amount of energy transferred from one side to the other. But what does the force times the length of the lever arm give us? This seems to be a critical quantity in a balanced lever system. If we shift the position of the weight on one side of the lever, the system is no longer balanced. So the values of ℓ_1 and ℓ_2 are quite important. The ratio, ℓ_1/ℓ_2 must equal the ratio w_2/w_1, and $\ell_1 \times w_1$ must equal $\ell_2 \times w_2$ if the system is in equilibrium. This product, the lever arm times the force acting on it, is commonly called the *moment of rotation* or the *torque*. The torque in one direction of rotation is equal to the torque in the opposite direction of rotation when a lever system is in balance.

A variety of lever systems have evolved in various species.

9–5 LEVERS AS SIMPLE MACHINES

A look at the distances traveled in relation to time in levers reveals an interesting characteristic. The rate of travel—the distance moved divided by the time in motion—is called *speed*. During the time our lever was tilting, one of the objects moved a greater distance than the other. It therefore had to have greater speed. The speeds of the two masses are inversely proportional to the forces. In other words, a large mass traveling slowly is balanced by a small mass traveling fast. We probably think first of using levers to increase the force we are able to exert, at the sacrifice of distance moved or speed of motion. Just as often, we use levers to produce a greater speed of motion. The bones of the body are themselves levers pivoted about the joints as fulcrums, with muscle tissue applying the force to the bone near the pivot point. The forearm is such a lever. The elbow is the fulcrum and the biceps exerts its pulling force at the point of attachment only a few centimeters from the fulcrum. This lever is not designed to amplify the force of the biceps. Instead, the force exerted at the hand is several times less than that exerted by the muscle. But the speed of motion of the hand is comparably greater than that of the connective tissue attaching the muscle to the bone. The same principle is seen in the even longer bones of the leg. [9] Is the speed at which an animal can run, as it uses the long levers of its legs to project itself forward, related to the lengths of the legs themselves? In your comparison consider two different members of the same species such as a greyhound and a terrier, as well as representatives of such different species as a race horse and a rabbit. [10] If the legs of a horse were twice as long, would it then be able to run twice as fast? Man has the muscular ability to travel faster under his own power if he can increase the distance moved

Skillful use of the entire body as a lever system.



The lever system of the arm produces speed rather than allowing us to exert a great force.

with each step. This is one of the advantages of a bicycle. $\boxed{11}$ Could a man learn to run faster on long stilts than he can with his feet directly on the ground?

A baseball pitcher puts great speed into a ball when he throws it by using his arm as a lever. The batter, if he can contact the ball near the end of the longer lever represented by his arms and bat together, can propel the ball much farther than it can be thrown. In such sports as jai-alai, golf, and tennis, the ball may be propelled at remarkable speeds. Primitive man discovered that he could project an arrow or spear to greater distances if he used a throwing stick, or atlatl. The housewife uses this principle in the flyswatter to eliminate the fly before he can dart away. In the mechanical typewriter or piano, short firm strokes of the keytops result in long swift strokes of the keys. Examine a model or a drawing of the human bone and muscle structure; look closely at the forearm in particular and its muscle system. Attempt to determine how many times greater the force exerted by the biceps must be than the load it supports in the hand. First, find the mass of the heaviest object you can support with your forearm horizontal. Next use the equation $w_1 \times \ell_1 = w_2 \times \ell_2$ to obtain the force exerted by the biceps. This force is also an estimate of the mass the biceps could support if attached directly to the lower end of the muscle.

An examination of models of human bone and muscle structure will show that it is made of many levers.

An astronaut in orbit experiences a sensation of weightlessness.

9–6 THE EARTH'S GRAVITATIONAL FORCE

The constant force we call gravity deserves more attention. First, it is constantly with us and thus taken for granted. Second, it is the only commonly occurring force that does not change as work is done against it. Because we live on a relatively large and uniform earth, the nature of gravity is unchanging. The laws of gravitational attraction as advanced by Isaac Newton predict that the attractive force of the earth should decrease as we move away from its center. This is measurable in changes of altitude from mountain top to valley floor, but the difference is so small that no method we have in our present study will detect it. This makes gravity a very useful force for our purposes. Unlike any system in which we store energy by doing work against the elastic forces, this force does not change as work is done with it or against it.

We cannot easily imagine how our life processes and routine practices would be changed if gravity were not present. This was a problem that had to be resolved before we could safely subject man or machine to the low gravity on the moon's surface or the apparent weightlessness of orbital travel. We are accustomed to the nearly constant pressure of our atmosphere; changes in pressure surprise us. In a similar way we are very aware of some new motion that makes it seem that gravity's force is either reduced or increased. The forces and sensations that we associate with the stopping or starting of a high-speed elevator, an airplane in turbulent weather, or a car traveling at high speed over a rolling road should remind us of how strongly we depend on a steady gravitational force.

These apparent changes in gravity can be measured and studied. When a force, acting either up or down, changes the motion of an object, it is either added to or subtracted from the constant gravitational force to produce the net force we experience. For example, as an elevator begins its descent, we, as passengers, begin to fall. The floor of the car no longer presses as strongly on us to support our bodies against gravity's attraction. If the elevator were falling freely, we would fall with it at exactly the same rate and would not require support. In an elevator, we would find that a spring or rubber band supporting a mass would change length during the period of starting or stopping, just as if we had changed the amount of mass attached to this same spring in our laboratory. [12] Why is this true? [13] Would the behavior of our platform balance be influenced by this kind of motion change?

9–7 ENERGY STORED IN ELASTIC MATERIALS

Thinking again about potential energy, we should realize that there are other methods of storing the ability to do work than by raising an object a distance against the force of gravity. There are other situations where we may exert a force, produce a change, and still be able to realize some return of the work done. One of the commonest is a spring. Springs are made of *elastic* substances. This means they have the ability to hold a certain amount of energy which they may release later. Under proper conditions, most matter shows some elastic properties. Let's study the elastic characteristics of wood.

9 4 For this experiment you will need a 30-cm length of string, a dowel 6 mm in diameter and 60 cm long, a meter stick, a support system for the dowel, and five 100-g masses. With the help of gravity we can apply known forces to the piece of wood and measure the changes that occur. Support the dowel between the two supports, one near each end. Tie the string in a loop around the dowel. Suspend one of the 100-g masses by the loop of string and center it on the dowel. Refer to the meter stick held behind the assembly to determine how much bending takes place as a result of the application of this force (the force is equal to gravity's pull on the mass—a 100 gf in this case). Record the height from the floor of the center point of the stick before and after applying the force. Now add another force by attaching a second 100-g mass at the same place and again read and record the height of the center point. Repeat this until a considerable bending occurs, but do not carry it so far that the

The force of gravity bends the dowel. As the force is released can the dowel return the work stored in its distortion?

dowel breaks. Now reverse the procedure, removing the masses one at a time and again reading for each force the distance from the floor of the point where the force is applied. Plot the *bending* of the dowel in millimeters as a function of the gram-force during both halves of the cycle.

$\boxed{14}$ What would be the effect on the experimental results if the dowel supports were moved closer or farther apart so that the effective length of the stick was changed?

Clamp the dowel to the top of a table so that one end extends out from the edge one-half of the distance that existed between supports in the first experiment. Add masses at the end of this beam in the manner you did before and again note the bending.

$\boxed{15}$ Would the length of time that the bending force is applied to a stick affect the tendency of the wood to return to its original position after the force is removed? Make a prediction and devise an experimental procedure to determine if your hypothesis is valid. $\boxed{16}$ Would the experimental results be different in the original test of the dowel if it had been soaked in water for 24 hours just before the test? Predict the results and check your hypothesis by experiment.

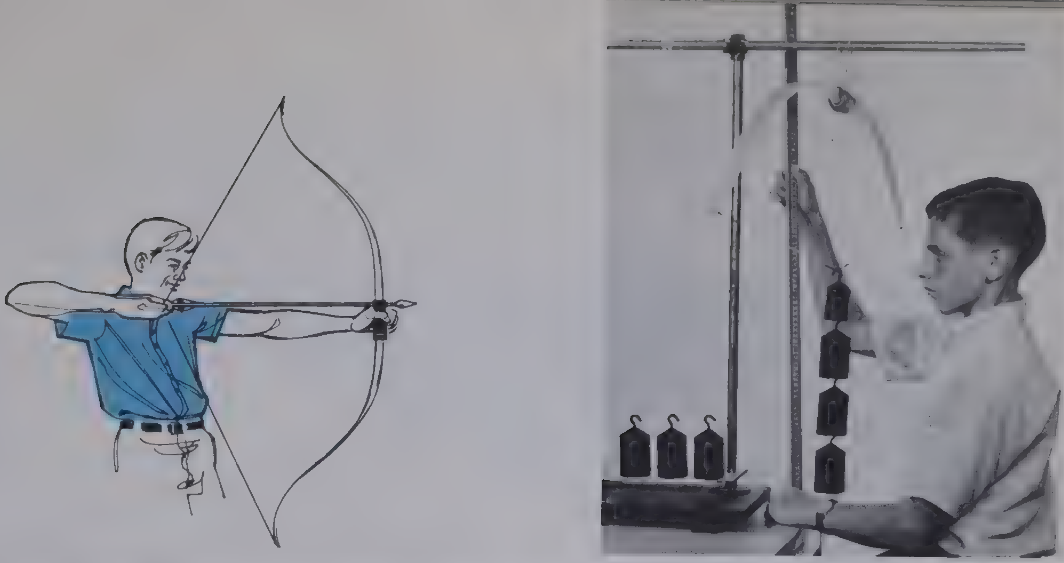
The spring balance we used earlier is valuable only because the spring inside the device is made of an elastic substance. Carefully examine the scale on this balance. Predict the shape of a graph that would show the stretch of the spring as a function of the magnitude of the distorting force. $\boxed{17}$ How does this shape compare with the graph you obtained while testing the wooden dowel?

What work can the rubber band do as the masses are removed?

9–8 PARTIALLY ELASTIC MATERIALS

9 5 Obtain a single wide rubber band (3–4 mm wide) and hang it by one end from a firm support. Attach a small known mass to the lower end of the loop and measure the length of the rubber band. One by one add equal masses of about 100 g each to the lower end of the loop. Record the readings of length with each added force. When you have gathered enough data, plot a graph of the stretch of the rubber band as a function of the applied gram-force. Next, remove the masses one by one and read the new length as quickly as possible after the force is reduced. Plot the new values of the elongation of the band on the same graph at the corresponding values of force. [18] How does this curve compare with the one made for the dowel? For a hint you may want to note the temperature change immediately after a rubber band has been stretched. Hold a rubber band by two opposite ends and pull it taut. Touch the center of the band to your lips or upper lip where your nerves are particularly sensitive to temperature differences. Now allow the band to return from its stretched position and again try to sense whether there is a temperature change.

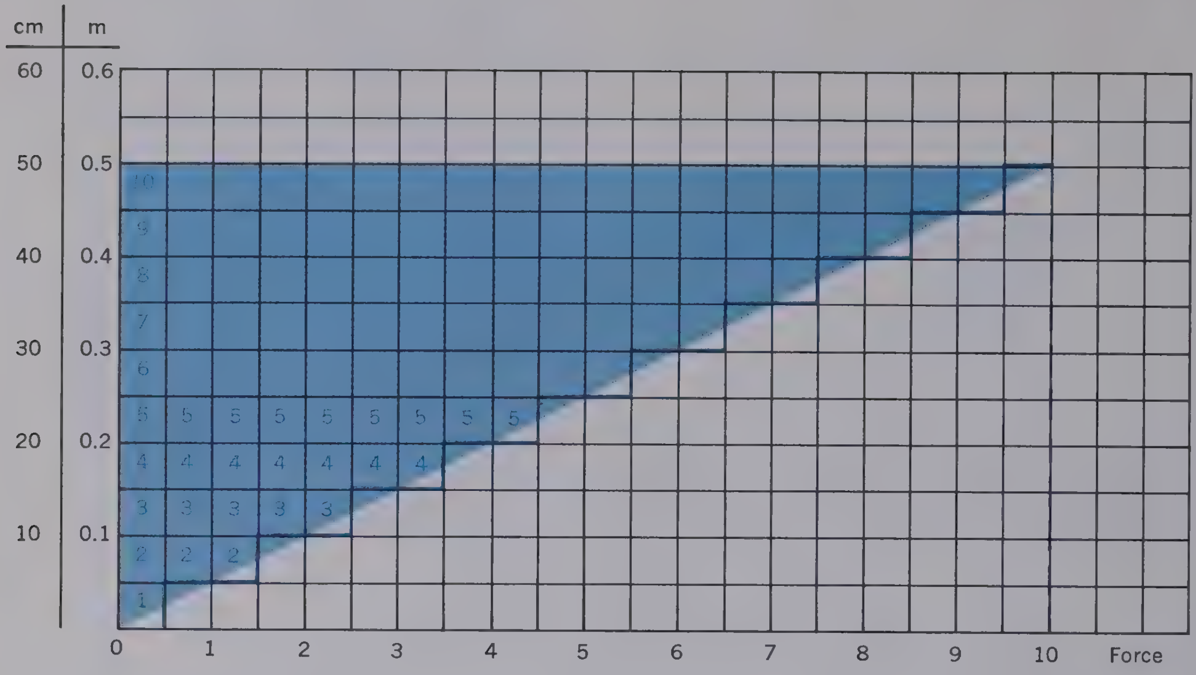
An understanding of the difference between the behavior of rubber and some other materials will give a better definition of the word *elastic*. In a perfectly elastic material, the energy that comes out is equal to the energy put in, provided the material is not deformed too much by the force doing the work. Materials that show almost perfect elastic proper-



The muscular strength of the archer provides the force to draw the bow. In the experiment at the right gravity provides the force. The amount of work done on the bow can be measured as it is gradually loaded with more and more mass and the force is increased.

ties under ordinary conditions include bamboo and glass. If these materials do not seem to you to be elastic it is because they have a low elastic limit. This means that only a little distortion may take place before the substance will no longer return to its original shape. The word *plastic* actually means "totally inelastic." Thus any force will permanently distort plastic substances such as mud or putty. Plasticity is a desirable property in a material that we wish to shape into a special form. Modern plastic materials are, in fact, plastic at the time they are being formed and later become hard and often quite stiff and elastic as they harden with cooling or chemical action. Some other common substances that, in our normal environment, seem to be quite plastic become hard and elastic if they are cooled to very low temperatures. Solder, a solution of tin in lead, is easily melted and bends or flattens with little force at room temperatures. At the temperature of liquid air, however, a coil of solder is an excellent spring, and shows extremely high elastic properties.

Let's look briefly at the work and energy situation in the application of elasticity to a simple machine such as the bow. When a bow is bent, a force is used and it moves through the distance required to draw back the string. Let's assume that on the graph shown on page 188 the straight line cutting through the blocks shows the relationship between the distance the arrow is drawn back and the force exerted. Whether this is actually the shape of such a curve (line) depends a great deal on the construction of the bow. [19] What amount of work has been done by the archer in drawing the bow? The force is not constant, but changes with the



amount of bending of the bow. For this reason we will analyze the problem by studying the force through various small distances and adding them together. Our analysis will not be exactly like the real situation, but the method is common in solving such problems.

Look at the graph on this page. Assume that it was necessary to use a 0.5 kgf before there was any motion of the string and that it then moved 5 cm. Again, when a total of 1.5 kgf was used, the bow string moved an additional 5 cm as shown in the graph. The force–distance curve would look like the step line. It would then be easy to calculate the total work in pulling the string any given distance. Let's consider drawing the arrow back to 50 cm. As the string moves the first 5 cm, the force is a constant 0.5 kg and the work is 2.5 kgf-cm. The area of the square marked "1" represents this amount of work on the graph. During the next 5 cm the force is a 1.5 kgf and this additional amount of work is 7.5 kgf-cm. The squares marked "2" represent this amount of work. With each addition of another kilogram force, another increment of squares is added. The sum of all of these small parts equals 250 kgf-cm. This is just half of the area bounded by 50 cm along the y-axis and 10-kgf units represented along the x-axis. The shaded area above the diagonal line is the same as the stepped area, 250 kgf-cm, and can thus be used to represent the work done. Using this technique it is possible to calculate with considerable accuracy the work done in stretching or bending any elastic object. In our bow and arrow example, the release of the arrow transforms into kinetic energy the potential energy that was stored up in the bending. [C, D]

9–9 ELASTIC PROPERTIES OF LIQUIDS AND GASES

Ordinarily we think of solids when we speak of elastic substances. [20] Do liquids and gases show any elastic properties? These forms of matter may not be bent or twisted because they are without definite shape. We do know, however, that the volume of a confined gas may be reduced by applying a force to it. Think, for instance, of what happens if you jump on an inflated inner tube. Any tight-fitting piston-cylinder combination will determine the relationship between the force compressing the gas and the volume that it assumes. [21] Can the volume of a liquid also be changed by applying a force? In arriving at an answer to this, think of how car brakes work, or how you feel when you do a "belly flop" in a swimming pool. [E]

9–10 IS ENERGY OF MOTION ALWAYS CONSERVED?

In our discussion we have postponed consideration of a significant fact. When a falling object strikes the floor, it does not seem to do as much work as the total energy of motion it possessed. During each successive swing, the arc of the pendulum grew smaller and eventually its motion stopped completely. Elastic materials, springs, springboards, or rubber bands seem to return reliably our input energy but after a short while an object bobbing at the end of a coil spring comes to rest. The balance or our meter-stick lever device will not swing up and down indefinitely; eventually the motion ceases. Our observations in everyday life tell us that kinetic energy seems to disappear. For example, a car rolling on a level road soon comes to rest, the driven golf ball rolls to a stop, and even the motion of billiard balls on a level table surface does not continue indefinitely. These occurrences are so common that it is not surprising that early Greek scientists concluded that it was the nature of matter here on earth (terrestrial matter) to stop, or come to rest. This seemed to be in contrast to celestial matter, which includes planets, the sun, the moon, and the stars; they appear to have continuing and unchanging motion, at least within the ability of man to measure any change in his lifetime. This was one of the misunderstandings that endured from the teachings of Aristotle and his students through the Dark Ages. Galileo, by carefully analyzing the problem and by doing experiments he thought would reveal the true situation, helped us to understand that there is no significant difference between the behavior of celestial matter and terrestrial matter and that both are subject to the same basic laws of the universe. This provided a basis for Newton, a century later, to formulate inductively his laws which predict the motion of all matter. Why does a chunk of matter continue plummeting through space while a similar piece of matter comes to a stop after moving over the earth's surface? Let's consider the pendulum in answering this question. As the pendulum bob swings, the air through which it moves exerts a small force against the matter. If you use a large object without much mass for a bob, you will note that the dying down occurs more rapidly than with a smaller denser object. [22] Why? [23] Where does the energy go? Does the air, once set in motion, continue to move out to the side, first to left and then to right, as the pendulum swings back and forth? If so, we should be able to detect small puffs of wind which have the same period as the pendulum. This certainly is not very evident. Is it possible that the energy is actually destroyed? Has it ceased to be, or is it possible that it still exists but in such small bits that we cannot detect it? If we look at a situation in which a great deal of energy of motion seems to disappear in a very short time, perhaps where it goes and what happens to it will be revealed.

9 6 You will need a block of wood and a piece of sandpaper. Rub the block of wood briskly with the piece of fine sandpaper. You are doing work, because you must exert a force in order to get the paper to move against the wood. [24] What happens to your hand and the wood surface as you continue this work? In rubbing the sandpaper against the wood, you may have found that it was nearly impossible to move the paper briskly, especially if you pressed the two surfaces together firmly. We call this retarding force *friction*. Friction is not evident unless another force is moving an object. Thus, friction always acts to reduce the motion of moving objects and, therefore, in the direction opposite to the motion. In this sense, the force that slows the motion of an object can be said to be negative with respect to the positive force that produced the motion in the opposite direction. The frictional force that stops the falling object as it hits the floor is not as easy to recognize as the frictional force that causes the pendulum ball to come to a stop.

While sanding the wood, you probably noticed that your hand and the wood got hot. We are inclined to say that heat is "generated" by friction. Did you notice that there is a very close relationship between

Sanding is a source of heat due to friction.

the amount of energy used and the amount of heat that seemed to be created? It is much more reasonable to believe that the heat in the wood and paper that causes the increase in temperature is simply another form of energy. [25] But what is the nature of this form of energy?

The radiometer is able to convert radiant energy to energy of motion. Heat generated by burning fuel is capable of producing motion in a bulldozer. Now we find that energy of motion may be converted quite simply to heat. Although scientists accept the particle theory in preference to the caloric fluid theory, we should examine the merits of both theories in relation to the conversion of motion energy to heat. [26] Can heat that comes from motion energy be explained by the caloric fluid theory, or does this heat confirm that the caloric fluid theory should be discarded? Without being able to isolate heat fluid, we might believe, for example, that it could flow out of our bodies and into the sandpaper to be spread over the wood. [27] When we stop rubbing the sandpaper, can we say that we stop sending heat fluid out of our bodies?

[28] Will the heat fluid explanation apply to energy transfer in the pendulum? To start the pendulum in its motion, we must pull the bob out to one side. This calls for lifting it slightly, hence, work is done. Energy in our bodies is used to do this work. Then, as the pendulum swings back and forth and finally comes to rest, it transfers energy to the air around it. We cannot detect an increase in temperature because the energy given up during each cycle is very small. But we know that meteors as well as rockets and other spacecraft are heated to melting and burning

temperatures by the frictional resistance of air. In the pendulum, the energy put into the bob by raising it to a starting position does not make any difference in the temperature. Therefore, none of the caloric fluid could have been stored in the object to be released later to the air through which it moved. It is relatively easy to assume, however, that the moving object gave up some of its motion to the air through which it traveled, much like a billiard ball may transfer some of its energy and motion to another ball with which it collides. If heat is related to the motion of particles, these experimental results help us improve our particle model. [F, G, H]

TAKING INVENTORY

As an object is lifted, work is done. Energy is used to counteract the force of gravity. As an object moves it has kinetic energy. In a new position an object can have greater energy than it had in its previous position. This position energy is called potential energy because it is capable of doing work, although at the moment it is not performing work.

Potential energy can exist in forms other than position. It can exist in a stretched spring, a drawn bow, or in compressed air. Once released, these devices can return this potential energy as a force that can do work.

As energy changes from one form to another it is conserved. This is a useful assumption borne out by evidence from experiments. In systems where energy seems to disappear we attempt to find out where it goes. Energy seems to take many different forms and the problem of accounting for it requires that we establish a link between all its various forms.

We have found that radiant energy, which usually produces a change in temperature, may also produce motion. Although we have not yet studied electricity, there are devices, commonly called solar cells, which convert radiant energy directly to electricity. This suggests that electricity is another form of energy. In Chapter 10 we will examine electricity and magnetism as an extension of our study of energy.

FURTHER EXPLORATIONS

A. If you lift a 1-kg mass a distance of 10 m and a 5-kg mass a distance of 2 m, have you done equal amounts of work? Use a graph to compare the work done in each case.

B. Make a pendulum with a toy balloon and string. How does it behave? Does it show a regular periodic motion similar to other pendulums you have observed? Try to explain its behavior. Can you make it behave differently by adding to its mass?

C. The particle theory has been introduced to you; you have studied much experimental evidence that indicates the theory is a good one. Consider the particles of matter and how they might behave to produce elastic and plastic conditions in matter. Include in your thinking the fact that solder is quite elastic at very low temperatures and that sulfur takes an intermediate plastic-elastic character when heated.

D. Use a bow and various gram masses to determine the amount of work done in drawing back the bowstring, following the method described earlier in this chapter.

E. Use the particle motion theory along with what you have learned about kinetic energy to explain the elastic qualities of gases.

F. As a motor boat travels through the water in a lake or river, does it cause any change in the temperature of the water? Explain your answer.

G. Use the particle motion theory to develop an explanation for friction.

H. Using past experiences and your new understanding of pendulums, try to analyze why the process of pumping a swing allows you to increase the amplitude. How does the pumping affect the amount of energy in the system composed of you and the swing?

FURTHER READING

ANDRADE, EDWARD N. DA COSTA, *Physics for the Modern World.* Barnes & Noble, Inc., New York, 1963.
An introduction to classical physics in narrative form. Includes simple explanations of the basic laws and principles of matter, motion, heat, electricity, and magnetism.

BLACKWOOD, PAUL E., *Push and Pull: The Story of Energy.* McGraw-Hill Book Company, New York, 1959.
A discussion of the nature of energy, its uses, and experiments you can do.

COHEN, I. BERNARD, *The Birth of a New Physics.* Doubleday & Company, Garden City, N.Y., 1960.
Traces the evolution of thought in physics from Aristotle to Newton.

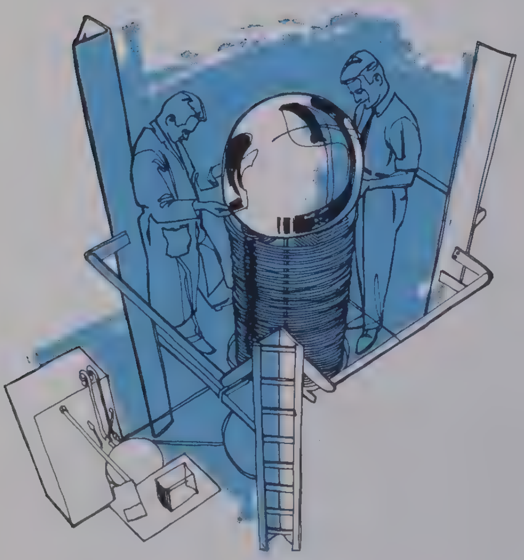
MARCUS, REBECCA B., *Galileo and Experimental Science.* Franklin Watts, Inc., New York, 1961.
An easy to read discussion of Galileo's contributions and experiments.

RUCHLIS, HY, *Orbit—A Picture Story of Force and Motion.* Harper & Row, Publishers, New York, 1961.
A discussion of inertia, acceleration, momentum, and gravity. Challenging problems are included at the end of the book.

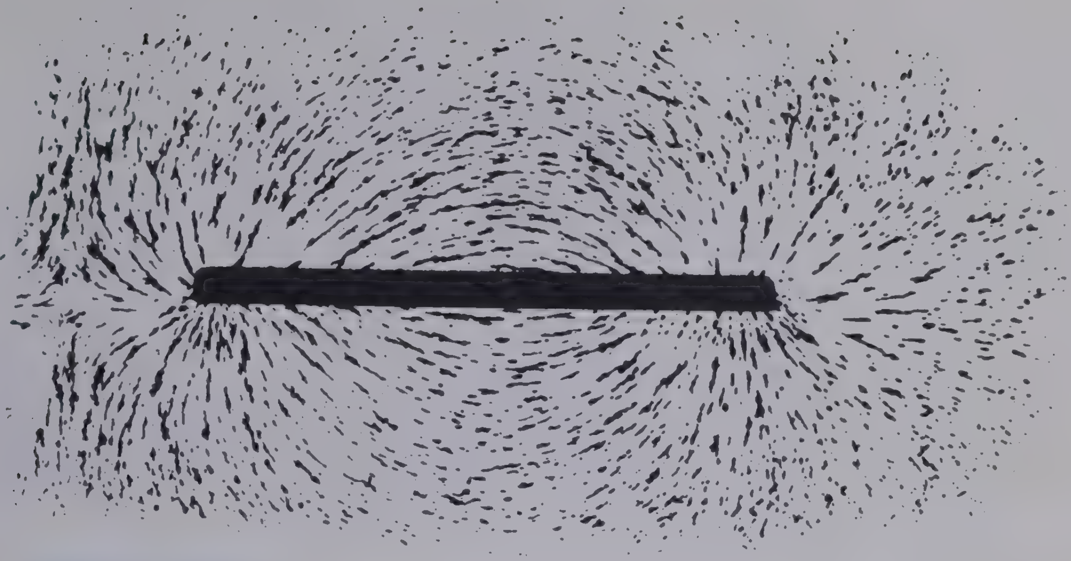
SOOTIN, HARRY, *Isaac Newton.* Julian Messner, Inc., New York, 1955.
A biography which explains Newton's work with light, gravitation, and motion.

10

Force Fields



10–1 Gravitation
10–2 Magnetism—A Force That Acts at a Distance
10–3 Magnetism and Electricity
10–4 Electric Charges and Fields
10–5 Modification of the Particle Theory
10–6 Electricity and Human Safety
10–7 Describing the Gravitational Field



10–1 GRAVITATION

We have been studying mechanical work in which a mass bends a stick or stretches a rubber band. One object in contact with another exerts force and uses energy. What is the source of this energy which the mass seems to possess? In studying the lever, we found that we can give a mass the ability to do work by raising it. A mass has greater potential energy in its elevated position than it has in the lower position. The real origin of this energy, however, is the attraction of the earth—gravity.

In Chapter 9 we said that early scientists thought terrestrial matter was basically different from celestial matter. These early scientists reasoned that bodies in space were destined to move in circles, whereas matter on earth had a natural direction of motion down toward the earth and tended to come to a stop. This was the explanation for why it is difficult to pick up a massive rock or why a thrown ball will fall. In the seventeenth century Isaac Newton advanced a single theory concerning all objects in the universe. This theory, which has withstood a variety of tests over more than 300 years, has become known as the *law of universal gravitation.* It states that every object in the universe exerts a force on every other object regardless of whether they are in contact. This force is directly related to the masses of the objects, and the ability to exert such a force is one of the properties of mass. The greater the mass of an object, the greater gravitational force it exerts. The law also states that gravitational force is related to the distance between objects. As the distance increases, the gravitational force decreases.

How does the strength of the magnetic field vary with the number of cards?

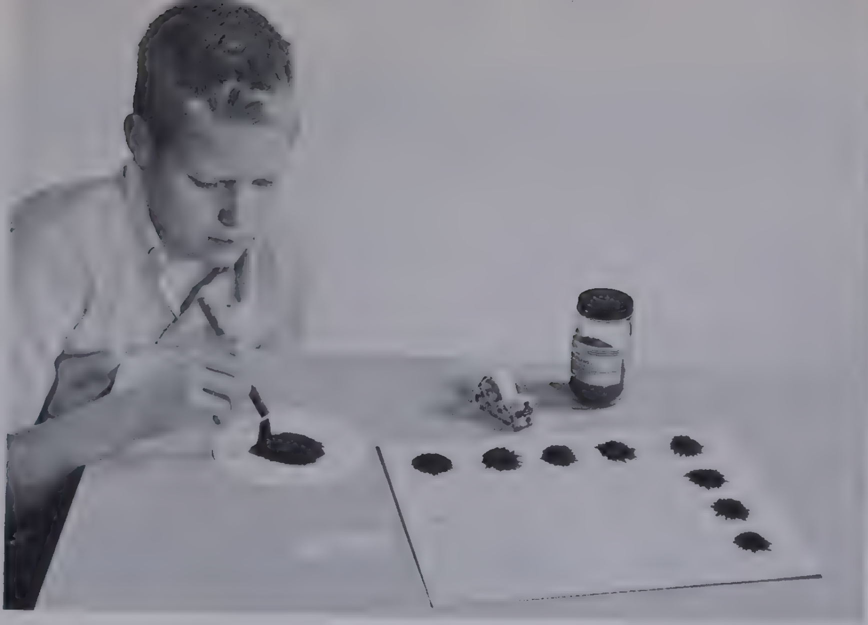
10–2 MAGNETISM—A FORCE THAT ACTS AT A DISTANCE

There is a force that acts at a distance that is not as general as gravity but follows some of the same rules. It is the force exerted by magnets. Let's look at magnets to try to gain more information about magnetic force.

Spend a little time studying magnets of various kinds and shapes along with a variety of other materials. Try to formulate some general rules regarding the force around them. Then classify the materials as to whether or not they seem to be influenced by a magnet. See if there seem to be any materials that keep the magnetic force from acting through them.

You have probably noticed that a magnet does not affect another object, say a nail, until it suddenly lifts the object all the way to the magnet. You may also have noticed that a particular magnet will do this when the magnet and the object are separated by the same distance each time you make a test. The force, in this case a force equal to the weight of the nail, seems to depend on the distance between the magnet and the object attracted.

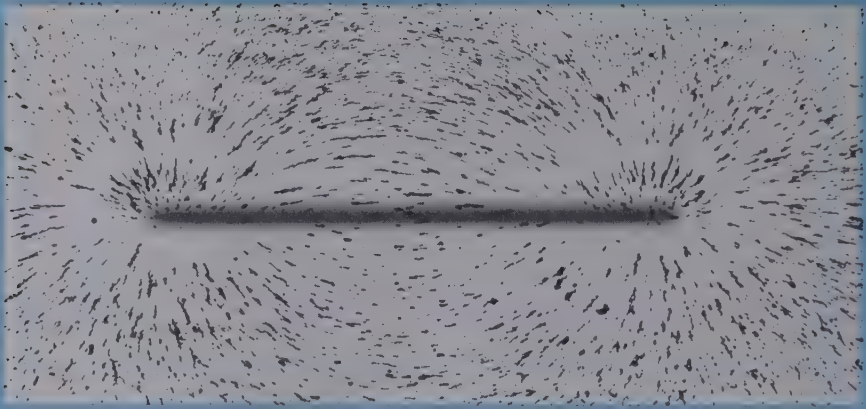
10 1 We can perform a rather crude test of the relationship between magnetic force and distance. You will need a long slender magnet that attracts objects at either end (the results with a short magnet or one that is U-shaped are not as good), several file cards 3 cm on each side, some adhesive



tape that sticks on both sides or spray tape, and 30-40 g of iron filings. Place a small strip of the adhesive (1-cm square) in the center on one side of each of ten of the 3-cm square cards. Press the end of the magnet against the tape on one card so that it sticks and then move it around slightly above the iron filings to pick up as many filings as possible. Lift the magnet, card, and attracted filings and place them on a sheet of paper. Pull the magnet from the adhesive so the filings on the other side of the card fall away into a small pile on the paper.

Again attach the end of the magnet to the tape of the card and then attach another card by its tape to the back of the first. You now have a card "sandwich" which has twice the thickness of the first card. Gather a second pile of filings using this sandwich. Keep in mind that the filings are two card thicknesses away from the magnet. [1] How does the pile of filings gathered with this arrangement compare with the first?

Repeat this procedure with three-card thicknesses, four-card thicknesses, and so on until you have reached ten-card thicknesses. Place each pile of filings lifted on a sheet of paper. What can you say about the strength (lifting ability) of a magnet as related to distance away from the objects it lifts? [2] About how many piles of the size lifted by the magnet with a 10-card thickness are contained in the pile lifted when only one card was used? A more accurate measure of the relative amount of material lifted would be to weigh each pile.

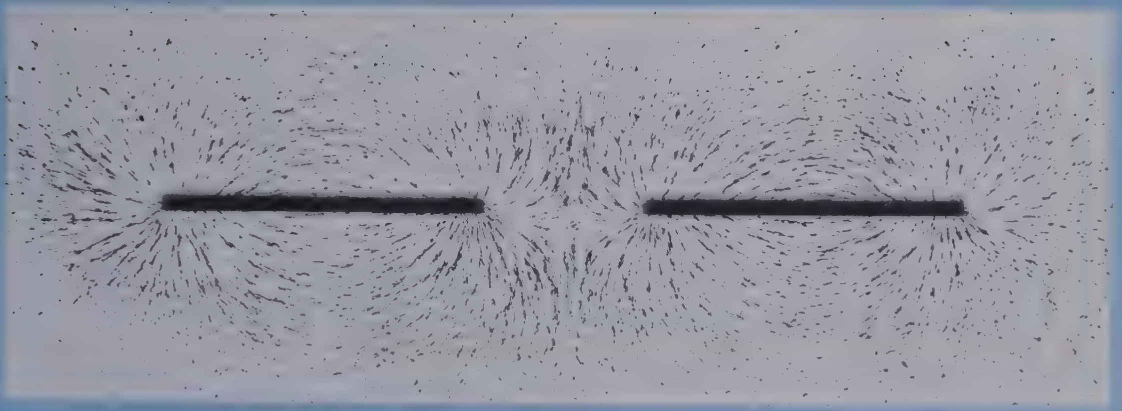


The use of iron filings is one way to indicate a magnetic field.

What can we learn about the structure of the magnet itself that will be of help to us? Does the magnet have a center of attraction? If so, where is it? Let's see if we can answer these questions by using the magnet and iron filings in a different way.

10 2 You will need the long, slender magnet and iron filings used in Experiment 10–1, a large sheet of heavy paper or lightweight cardboard, and a spray can of clear varnish. Put the magnet flat on the table and cover it with the heavy paper so that no part of the magnet is less than 10 cm from the paper's edge. You may want to support the edges of the paper with pencils, a glass rod, or some other nonmetallic material so it will be level and flat. Now sprinkle a small amount of filings on the paper and tap the table gently so that the filings take up a pattern. The distribution may not be uniform, but at least three-fourths of the paper's surface should be visible between the filings. Hold the spray can 30 cm or more from the paper and spray an even coating of varnish over the filings. When the varnish has dried the pattern stays in place and you may study it more easily, even after the paper has been removed from the magnet. As you examine the pattern, note its three-dimensional character near where the end of the magnet was. Sketch the pattern in your laboratory record book.

The ends of the magnet from which the lines of filings seem to point are called *poles*. Very early in the history of the study of magnets it was learned that magnets pointed in a generally north-south direction. The



Are the poles of the bar magnets alike or unlike?

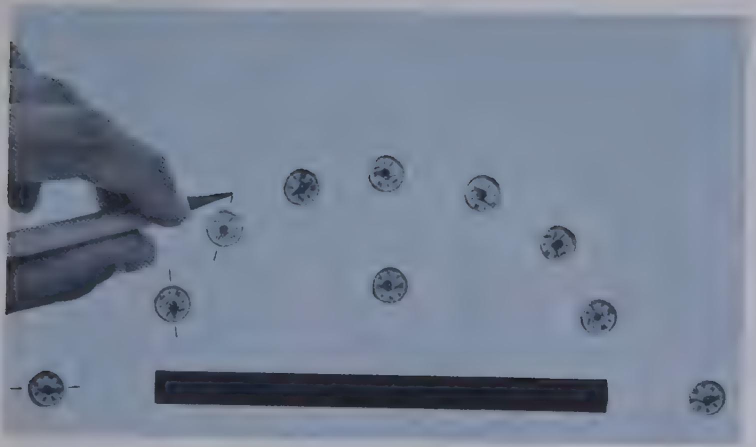
poles were named "north" and "south," depending on the direction they took. The region around the magnet is called the *field.* We have settled on a custom of referring to the direction of a magnetic field as the direction in which the north pole of a compass points in that field. The lines formed by the filings are called *field lines.* Actually there are no such lines until some magnetic material is present to point their direction. Field lines show the path an attracted object will tend to follow as it moves toward the magnet.

After you have studied the pattern of an entire force field around a bar magnet, place two such magnets on a table with their unlike poles about 2 cm apart. Put a piece of lightweight cardboard or unruled paper over them and sprinkle iron filings on the paper. This will show the pattern of the force field between unlike magnetic poles. Either spray this pattern to preserve it or sketch it in your record book. Now place two like poles of a bar magnet about 2 cm apart and repeat the process of obtaining a pattern. 3 How do the two patterns compare? 4 Propose a hypothesis to explain your observations.

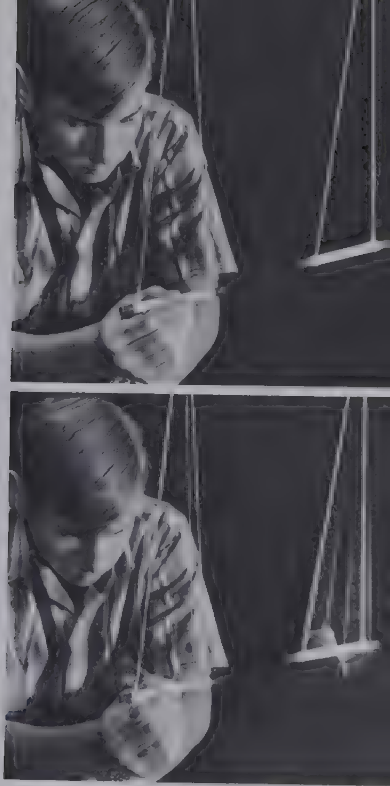
Using other magnets of different shapes, make a variety of patterns by arranging the magnets in different ways with respect to each other. Characteristics of magnetic fields can be further explored by placing iron nails or bolts in the force fields to determine what effect these have on the lines of force. 5 Recognizing that large iron objects may distort the pattern, do you think the tiny iron filings might affect the force field?

Cast another pattern of a force field of a strong bar magnet. This time spray the pattern with a fast-drying paint. Pour off the filings so you can have a smooth paper on which to work. Draw pencil lines toward the pole centers along each of the field lines at one end of the magnet. $\boxed{6}$ How does the spacing of the lines along the axis of the magnet change with distance out from the pole center? $\boxed{7}$ Is the spacing of the lines related to the force the magnet exerts at these points? Devise an experiment to measure this force. $\boxed{8}$ Is the force the same at the side of the magnet where the lines are equally spaced?

On this picture of the field pattern, draw a continuous line which intersects each field line at right angles. $\boxed{9}$ What can you predict about the force at various points on this line? $\boxed{10}$ Are the field lines equally spaced along this line? $\boxed{11}$ Would the same amount of work be required to move a small piece of iron from the pole center to any point on this line? The line is called an *equipotential line* because every point on it has equal potential magnetic energy. Draw a few equipotential lines in the patterns of magnets of other shapes. $\boxed{12}$ What relationships can you find between the shape of a magnet and its ability to lift objects?



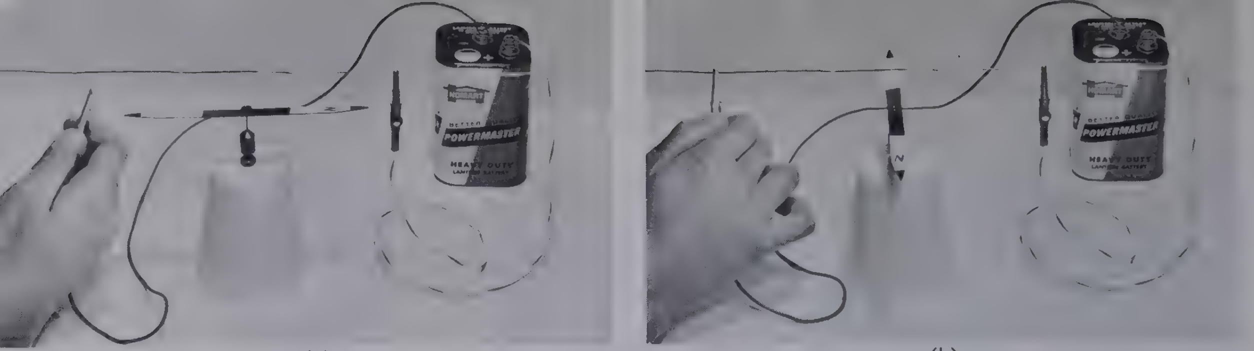
Using compasses is another way to trace the lines of force of a magnetic field.



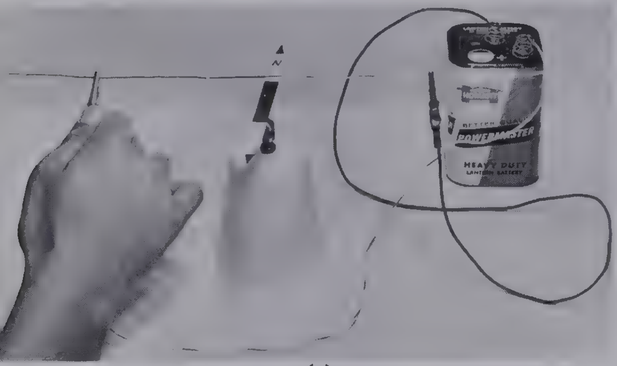
The interaction of magnetic fields.

Obtain a small compass (1 cm in diameter) and place it beside one pole of a magnet. Do this on the center of a page in your laboratory record book. Make a pencil dot on the paper at the end of the compass needle. Move the compass out until the opposite end of the needle is at the dot and place another mark at the remote end, as shown in the illustration. Continue to move the compass and mark the needle position until it again touches the magnet. One definition of a magnetic field line describes it as showing, at any place, the direction in which a compass will point. [13] Do field lines made by joining these dots in succession have the same general shape as those made with filings? A compass needle is a small magnet. [14] Does it distort the field of a magnet?

Fasten a 30-cm length of thread near *each* end of a bar magnet. Tape these two threads to a meter stick supported horizontally between two chairs. The magnet should be level. Attach another similar magnet in the same way, so that the near ends of the two magnets are about 10 cm apart and along the same line. Starting with both magnets at rest, twist one so that it oscillates in a horizontal plane. Notice what happens to both magnets over a period of several minutes and propose a hypothesis to explain your observations. [A, B]



(a) (b)



(c)

In (a) there is a break in the circuit; note the position of the compass needle. In (b) the circuit is complete. In (c) the circuit is complete but the current flow is reversed.

10–3 MAGNETISM AND ELECTRICITY

10 5 For another experiment with magnetism you will need a coil of wire containing several hundred turns of no. 28 copper magnet wire, a flashlight dry cell, and a small compass. Connect the coil of wire to the terminals of the flashlight cell and explore around the coil with the compass. Make a sketch of the apparatus in your record book, showing the direction of the field (the direction pointed by the compass needle) and the positive and negative ends of the dry cell. What influence does the manner in which this coil is connected have upon the field? Reverse the electrical connections and again explore the field. Plot such a field or devise a method of casting a pattern with filings to provide a record of the characteristics of this field. [15] What is the effect of placing materials such as iron, aluminum, wood, or glass in the center of the coil? [16] Does the shape of the field change if the coils are bunched close together as compared to farther apart?

Although we have not studied electrical circuits, there seems to be a link between electrical energy, which can produce radiant energy in a light bulb or kinetic energy in an electric motor, and a magnetic force field. The nature of electrical energy is so vague that hundreds of experiments and years of theorizing were necessary just to bring about our present incomplete understanding.

10–4 ELECTRIC CHARGES AND FIELDS

It has been known for thousands of years that certain dry substances behave strangely when rubbed. One of these substances is amber, a hardened resin from ancient plants. Modern technology has provided a host of other materials that have similar properties and are easier to obtain. Let's conduct tests with a few of these materials to learn about electrical fields.

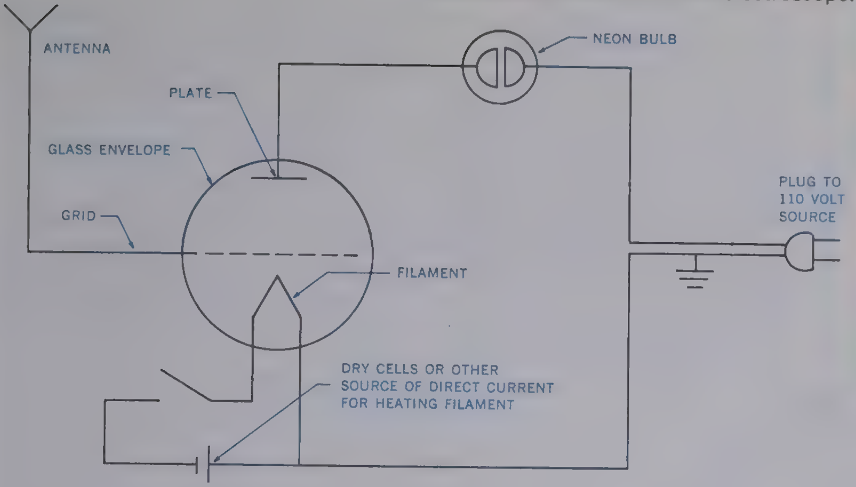
Cut from a plastic garment bag several loops 5–10 cm wide. Hold one as a double strip and draw it back and forth across wool or fur several times. Turn it over and rub the other side. Let it hang from one end and observe the behavior of the rubbed sides of the loop. Bring it near your other hand, your clothes, furniture, walls, a water faucet, and other objects around you. [17] Does it respond in a similar manner when you bring it near another plastic strip that has been rubbed the same way? [18] How does it react to a sheet of the plastic that has not been rubbed? [19] Is there evidence of a field acting at a distance? 10 6

The plastic film behaves differently than it did before it was rubbed; it is said to be electrically charged, and we consider it to be surrounded by an electric field. Our results with the plastic strips indicate that electric fields are strong and operate at a considerable distance. Try to determine if the field resembles that of a magnet. [20] Does it appear to interact with a magnetic field?

To investigate electric charges and fields still further, tie a length of sewing thread to the neck of an inflated balloon and hang it from a support. It should not be able to touch the support as it swings. Make an identifying mark on one side of the balloon and then rub that side with fur, wool, or hair. $\boxed{21}$ Does the rubbed surface show any evidence of being charged? As the balloon hangs, it is free to turn or swing in response to the presence of electric fields. It can help you determine the nature of these fields and some of the electrical properties of matter. Used in this manner, the balloon provides evidence of the presence of electric charges and may be called an *electroscope*.

Bring the charged balloon near your hand, the wall, or the water faucet as you did the rubbed strip of plastic bag. $\boxed{22}$ Can you suggest a possible model for electric charge that would give an explanation for the behavior of the balloon? Like some materials in a magnetic field, some objects seem to be either attracted or repelled in an electric field. There are many substances that produce an electric field when rubbed, especially if atmospheric conditions are proper. Rub and test different substances. $\boxed{23}$ Do all of the objects you rub and test appear to repel the charged face of the balloon?

Once again, let's try to construct a mental picture to discover the cause of the behavior of the materials we have been rubbing. This cause is what we ordinarily call electricity. Since the change in the material seems to come about with rubbing, we might assume that we are rubbing something off of one object and onto another. $\boxed{24}$ Does this seem to make



Schematic drawing of a vacuum tube electroscope.

The interior of a vacuum tube electroscope.

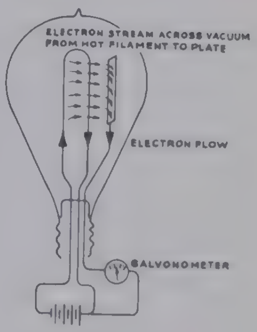
any difference in the color or in the weight of either object? In our study of heat we considered the possibility of a massless invisible fluid which causes changes in an object. While we did not find any direct evidence of this fluid, it served to help us explain a number of our observations. Would such a massless, invisible fluid help us build a theory of electric fields?

Using only the balloon as an indicator, it is difficult to conduct meaningful tests of our theories, especially if we realize that some of the effect of rubbing may disappear with time. Instead of the balloon, we can use a device called a vacuum tube electroscope, which detects electric charges just as the balloon electroscope does, but with greater sensitivity.

The significant parts of the vacuum tube for our purposes are the *cathode* (or *emitter*), the *control grid*, and the *plate*. These three elements, enclosed in the glass bulb, are in a region of such low pressure that we refer to it as a vacuum. Only a small fraction of the air remains to interfere with the motion of the particles in which we are now interested.

Thomas Edison was the first man to note the effect that is used in the vacuum tube. During his early work with the incandescent lamp in 1883, Edison was trying to find a way to lengthen the life of the filaments. He noticed that a current would flow in a circuit outside the lamp if he connected one terminal of the current source to the filament and the other to an electric terminal that extended through the glass bulb as shown at right. In other words, current flowed across a wireless gap. Edison's first incandescent lamps were emptied of air so the filament would

EDISON EFFECT



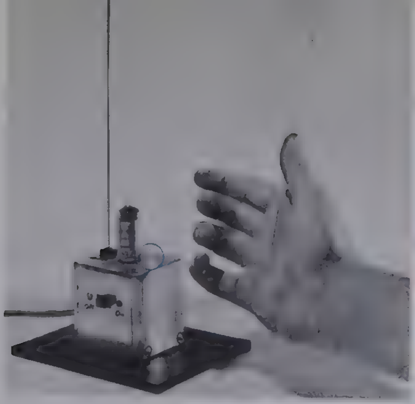
not burn. Therefore, he was really working with a vacuum tube. To Edison, these observations were interesting but not important in solving the problem he was working on. He recorded his observations and discontinued this line of investigation, since the unusual effect did not seem to influence the length of time his filament would endure.

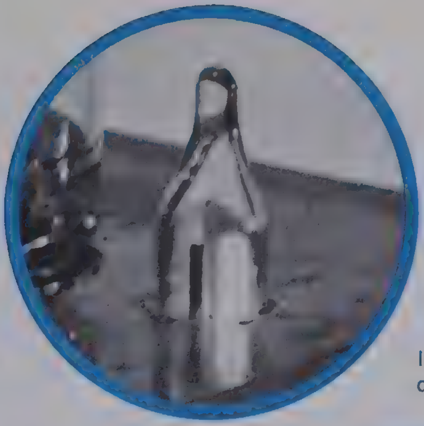
The importance of this effect, now known as the *Edison effect*, was not realized for many years. That the entire electronics industry is built around the vacuum tube indicates the importance of Edison's discovery. Today our explanation of the Edison effect is basic to electrical theory. The heated filament in the lamp or in the vacuum tube becomes the source of countless numbers of tiny charged particles we call *electrons*. The heat causes the rate of vibration of the particles of the metal in the filament to become so great that some of the electrons are "boiled off" the metal. When the plate which extends through the glass bulb is connected to the positive terminal of the battery, these "boiled-off" electrons move through the empty space between the plate and the filament and a current flows.

Our observations with rubbed objects indicated that *like charges repel* each other (two similarly charged objects tried to separate) and that *unlike charges attract*. One of Edison's discoveries was that the tiny current flowed in the lamp only when the filament was connected to the negative terminal. Thus we can deduce that the tiny charged particles possess a negative charge.

What does this mean in terms of the charges we found by rubbing? Is the rubbed object positive or negative? Our electroscope that utilizes a vacuum tube will show something of this. Notice that there is one added element, the *grid*, in our vacuum tube that was not in the Edison lamp. If we can apply a charge to this grid, it will have a field that will influence the transfer of charges through the vacuum. If we make it positive, the grid will speed charges away from the filament and more of them will pass to the positive terminal. If the grid is made negative, it will retard passage of charges from the filament to the positive terminal. The neon lamp that is part of the circuit is a sensitive indicator of the flow of current to the positively charged terminal, the plate.

10 7 Connect a small 1.5-volt dry cell between the filament and the grid. Notice that the lamp becomes brighter when the grid is made positive as compared to the opposite connection. This confirms that the charges are negative. Now rub the plastic strip again with fur or wool and bring this plastic *near* a wire connected to the grid only. [25] From the change in brightness of the lamp, would you guess that the rubbed plastic has a positive or a negative charge? Repeat with a rubber balloon. [26] Is there any material you can find that shows evidence of having an opposite charge when it is rubbed?

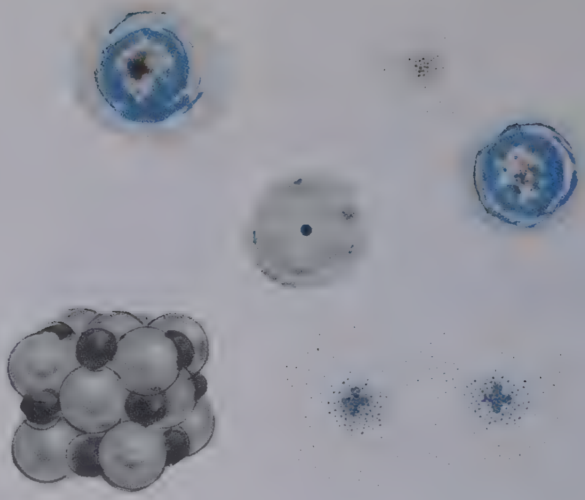




Is there a positive or negative charge on the plastic strip?

This similarity of the effect of charged substances and dry cells on the vacuum tube is one small link between the subject we call *static electricity* and that called current electricity. The first such link was found by Benjamin Franklin in his classic experiment with the kite. His experiment was dangerous; it is almost miraculous that he was not killed by the lightning. But his conclusions were sound, even though his experimental method was not wise. Franklin helped us understand that the events represented by a lightning discharge entail the same events as the spark discharge one experiences after walking across a rug or sliding across plastic seat covers. A spark discharge, we have shown, is the same kind of phenomenon as that which can result from the chemical action inside the cells of a battery.

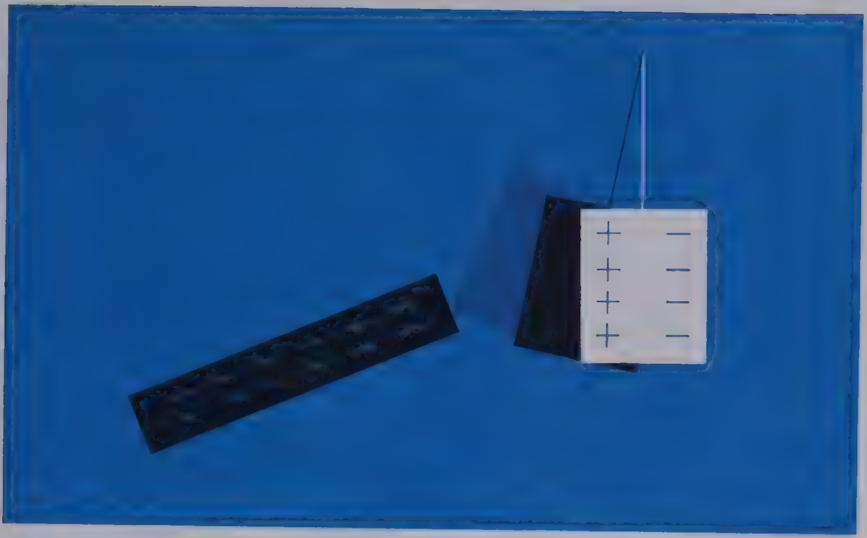
Following the custom originated by Franklin, we give the name *negative* to the charge concentration resulting from an excess of electrons and *positive* to a deficiency of electrons. An uncharged object, then, has neither a shortage nor a surplus of these charges. We believe that all matter ordinarily exists in this neutral state. [C, D, E]



Many of our observations about matter are explained by a concept of atoms that are themselves made of particles.

10–5 MODIFICATION OF THE PARTICLE THEORY

In our earliest experiments we set out to determine the structure of matter. From many observations of many experiments, we decided that a particulate model for matter was the best. In a given substance, the particles are more alike than grains of sand, or even the marbles with which we made our crystal models. To develop a clearer picture of electricity, we must make another assumption regarding these particles of matter. They, in turn, must be composed of even smaller particles. One of these particles, the electron, has a property we call *charge* which carries with it a field capable of influencing other charges at a distance. The fact that neutral particles of matter exist gives us reason to suspect that there are also positive particles corresponding to the electrons which, when taken in equal numbers with the electrons in any particles of matter, show no total charge. These ideas, then, are an extension of the particle theory of matter. Another name for the particle theory is the *atomic theory*. Thousands of experiments have been done that supply the basis for a theory of the nature of atoms. These experiments do show that atoms, the fundamental particles of our particle theory, are complex organizations of electrons and other electrical particles. The mass of the atom is concentrated in a *nucleus* but separated from it are the negatively charged electrons. The nucleus contains particles called *protons* which are positively charged. Every kind of matter we encounter has its unique number of protons and a corresponding number of electrons. The Edison effect indicates that some electrons can be separated from atoms. We do not



Charging a piece of aluminum foil by induction.

find similar evidence for positive charges separating from the atom. Instead we are led to assume that those particles with positive charge must be an inseparable part of atoms. When the positive charges move, the atom moves.

The freedom with which electrons move from one atom to the next, as envisioned in an atomic theory for matter, permits us to classify most materials as being conductors or insulators of electricity. Electrical conduction is the process of transfer of electrons. The materials we used to generate charge concentrations, such as the plastic or the rubber in the balloon, show little tendency for this transfer. The charges we established by rubbing did not escape to our hands or to other objects, even if we may have touched the plastic or rubber. There may have been a slight local discharge, but this did not eliminate all of the charge on the balloon. At the other extreme, in that common group of substances we call metals, the transfer of charges seems to take place very readily. Just as we can take advantage of the insulating characteristics of the plastics or of rubber to establish a charge by rubbing, we can utilize the conducting characteristics of metals to provide them with a charge by a process we call *induction*.

Obtain a 20-cm square of aluminum foil and suspend it by a thread held in your hand. Hold it near a piece of charged plastic. Keep the foil from touching the plastic by holding it with the other hand. Release your hold on the foil and then pull it away with the thread. Now test the foil for charge with the vacuum tube electroscope. Put a smaller strip of foil suspended by a thread near the large sheet of foil. Observe the

Voltages of some common electrical systems.

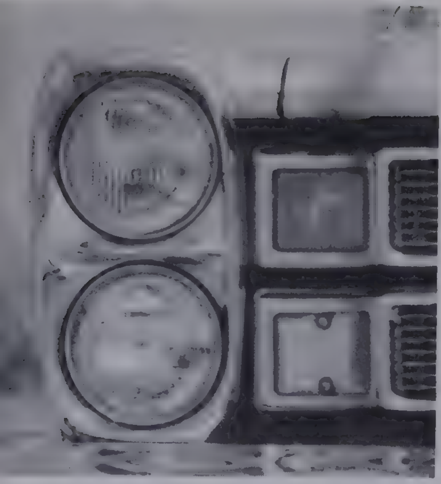
Incandescent light — 110 volts.

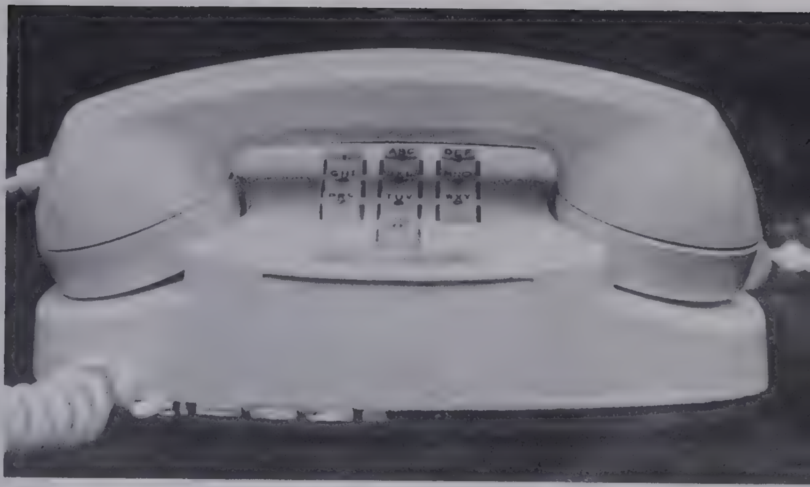
Transistor radio — 9 volts.

behavior of the small strip carefully. Then see if the small strip is now attracted to a charged balloon. You have charged the sheet of foil by *induction*. While you were touching the foil in the field of the negatively charged plastic, negative charges (electrons) passed (were induced) away from the plastic, through the foil, and into your body. Removing your hand before you drew the foil from the field left a deficiency of negative charges on the metal. This is to say it had a positive charge.

Whether matter is charged directly or by induction, some energy is stored and an electrical field results. Remember that stored energy is also called potential energy. We can describe potential electrical energy as the amount of work done to charge an object, or as the amount of work the charged object may do as it returns to a neutral condition. Instead of considering, as we did with gravitation, the weight of the object and the distance through which it moves, we refer to the number of charges and the *potential difference* between two points. The unit of measurement of potential difference is the *volt*. The voltage between two points is a measure of the amount of work that must be done to move charges between these points. The quantities are not as easy to measure directly as weight and distance. We need to use an electric meter designed to interpret the indirect evidence it receives into the units of energy we have described.

The range of electrical potentials runs from some much smaller than that produced by a common dry cell to the millions of volts that result in a lightning discharge. Tremendous numbers of charges in clouds are concentrated as the complex processes of a building storm separate some of the electrons from their atoms in the air and water. The sudden discharge of this concentration by a bolt of lightning spends most of the


An automobile — 6 or 12 volts.


Telephone — 48 volts.

energy, resulting in light and heat and other changes we shall consider in Chapter 11.

In the cell used to power a radio or light the bulb of a flashlight, the total potential energy is only 1.5 volts, but chemical action sustains this ability to do work over a long period of time. Finally, when the cell can no longer do as much useful work, the potential drops, and we find by connecting an electric meter across the terminals of this cell that the voltage is lower than a fresh cell. Electric meters are not equipped to measure such sudden bursts of energy as those during a static discharge but we can get a clue from the length of spark that results. If the spark from your body after you have walked across a rug is about 1 cm in length, you established a potential difference of 20,000 volts or more between your body and the earth. [F]

10–6 ELECTRICITY AND HUMAN SAFETY

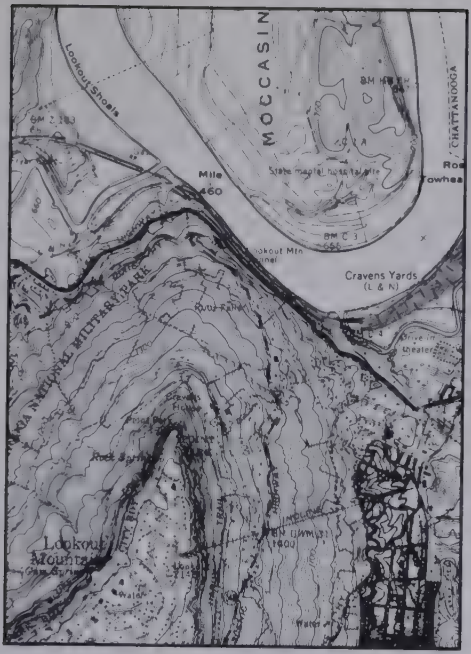
The numerous nerves found in every part of the human body are conductors of electricity. In fact, the impulses that travel along nerves to produce sensations of pain, temperature, sound, taste, and smell are electric in nature. So are the impulses that cause muscles to contract. The brain and other parts of the nervous system generate electric pulses at a rather regular frequency. These may be detected, recorded, and studied. An electrocardiogram (EKG) is the record of the electric pulses accompanying the beating of the heart; an electroencephalogram (EEG) shows impulses generated by the brain.

The amount of current needed to light a tiny lamp is more than enough to make any muscle in the body contract. But the low energy of a 1.5-volt dry cell and the static charges built up by rubbing cannot produce such a current through normal body tissue. High voltages are dangerous because they are capable of doing a large amount of work, that is, they can cause a large current of electric charges to flow. The contact one usually makes through dry skin with wires on household voltage (120 volts) is considered safe. But a person in a tub of bath water, exposing a large surface area to conducting water, would be killed if he contacted a wire connected to a 120-volt source. A deadly current produces a spasm in human muscle. The diaphragm convulses and the victim exhales violently. The heart muscles contract and the heartbeat stops. Less than a minute after the brain is deprived of fresh blood, it begins to deteriorate. Unconsciousness and death soon follow.

10–7 DESCRIBING THE GRAVITATIONAL FIELD

We should not leave the study of force fields without giving more attention to gravity. We cannot increase or decrease the force of gravity as we have done with electric and magnetic forces, but we can plot some of the field lines. As with magnets, the lines are the paths of objects moving in the field. The gravitational field of the earth, then, is composed of all the paths an object might follow if it were released from a rest position to fall to the earth. These lines are called *plumb lines*, or vertical lines. Our words "up" and "down" refer to the two directions along these lines. Equipotential lines may be drawn across the earth's gravitational field just as they may be drawn across a magnetic field. For example, if a line were drawn across the smooth surface of the ocean it would be equipotential. That is, the potential gravitational force would be the same at any point on it. A *contour map* shows lines joining equipotential points on the surface of the earth. Examine the map shown in the illustration.

The earth is not a perfect sphere. It is slightly less spherical than a billiard ball and considerably more spherical than an orange. The best description of the earth's shape has been gained from the orbits of our artificial satellites. If the earth were a perfect sphere with gravitational field lines that were the true radii, so that all lines passed through a common center point, the orbits would be more regular than they are. Irregularities in the field pattern of the earth have also been used to locate abnormally large underground masses of rock or oil that may have densities different from their surroundings and therefore be a source of variation in the gravitational field of the earth.



Segment of a contour map showing Lookout Mountain in Tennessee.

The earth is surrounded by a magnetic field as well as a gravitational field because the earth itself acts as a gigantic magnet. This is why a compass, which is a small magnet, may be used to determine geographical direction.

TAKING INVENTORY

We have learned that the region around a magnet is called its force field. The magnetic energy represented by this force field can do work on some substances. Work must be done to remove these same substances from the field of a magnet. Electric currents produce a magnetic field around the wire through which they flow.

Neutral objects can be electrically charged by rubbing. Negative electrical particles are called electrons. Bodies with an excess number of electrons become negatively charged, whereas objects with fewer than the normal number of electrons are positively charged.

The earth has both magnetic and gravitational force fields around it. Compass needles respond to the earth's magnetic field and all objects on the earth respond to its gravitational force field.

This chapter added electrical and magnetic energy to the forms of energy we have studied. It has brought our attention to the existence of force fields.

To make our study of energy complete, we need to examine it in chemical form. Chapter 11 is concerned with chemical energy and its relationship to the other forms of energy we have studied.

FURTHER EXPLORATIONS

A. Are iron filings already magnets before they are brought into the field of a bar magnet? How can you find out?

B. As you examine the pattern of iron filings in a magnetic field, you will notice that the lines of force converge on a point that is not at the surface of the magnet. Try to measure the distance from the surface of the magnet to this point. Is your original generalization about the strength of a field as a function of distance from the pole center correct? If not, can it be modified?

C. In Section 10–4, we stated that atmospheric conditions affect whether materials will produce an electric field if they are rubbed. Devise an experiment to test this statement.

D. Using a charged rubber balloon as an electroscope, determine the polarity you have from the static you gather while walking across a thick carpet. Is it the same charge you would get from sliding on plastic upholstery?

E. Use your knowledge of the particle theory to describe the nature of conductors and insulators. Can electrons move easily in some substances and not in others? Why can a charge be collected on some insulating materials so easily? Are pieces of plastic and rubber balloons examples? Or are these things partially conductive? Could charges build up if all matter was quite conductive?

F. Describe the manner in which electric charges were transferred, and the result, when the strip of foil was touched against the larger sheet of foil. Does your description explain what happens to the piece of foil?

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Contains instructions for several experiments using simple equipment.

11 Chemical Energy

11-1 Energy, Ions, and Chemical Reactions
11-2 A Chemical Cell
11-3 The Conductivity of Solutions
11-4 Chemical Cells, Conductivity, and Ions
11-5 Chemical and Physical Changes Compared
11-6 Energy in Chemical Reactions
11-7 Chemical Reactions of Nonionic Substances
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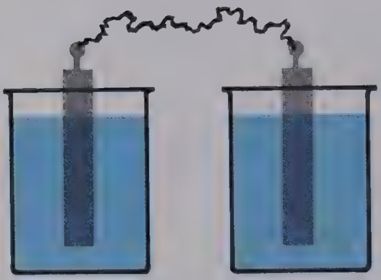
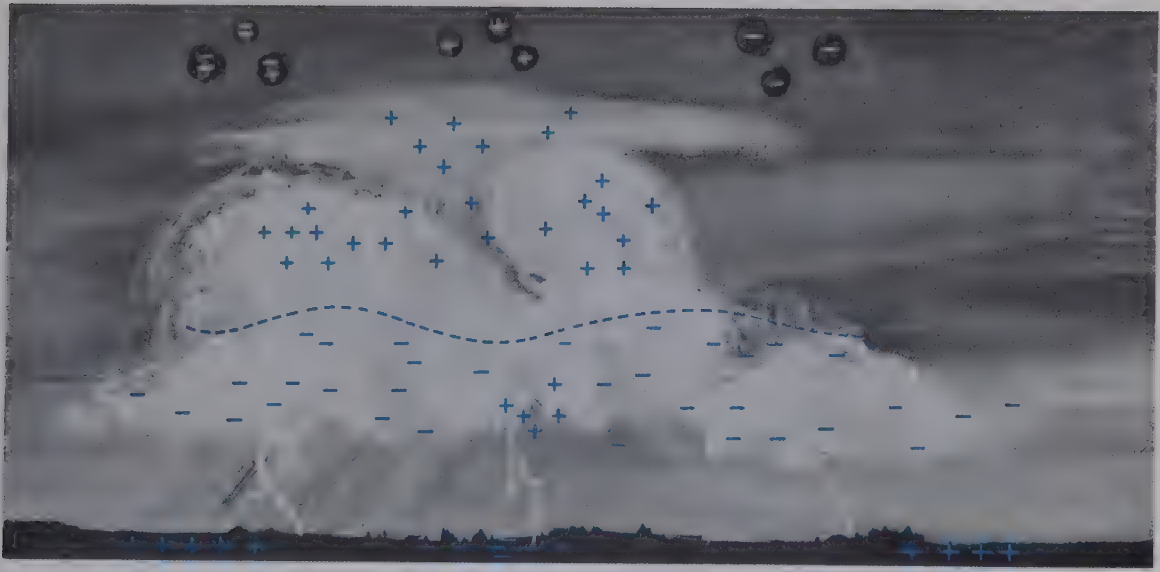
11-9 Reaction Rate and Concentration
11-10 Reaction Rate and Surface Area
11-11 Reaction Rate and Catalysts
11-12 Reaction Rate and Electrical Energy
11-13 Chemical Equilibrium
11-14 Why Do Chemical Reactions Occur?





Accumulation of electric charge according to the particle theory. Lightning is a natural electric spark.

11–1 ENERGY, IONS, AND CHEMICAL REACTIONS

Let's consider a thunderstorm. On a summer afternoon, menacing storm clouds gather. The day grows dark and the quiet, moist air is oppressive. Suddenly, the air is moved by violent gusts of wind. Large drops of water splash heavily and the sky is split by a brilliant flash of light. This is quickly followed by the crashing sound of thunder. We may understand the events of a thunderstorm, but how carefully have we considered what is happening? If lightning strikes a building, the changes are quite evident, but there are more subtle, more significant changes of energy which should be examined as part of this rather stupendous event.

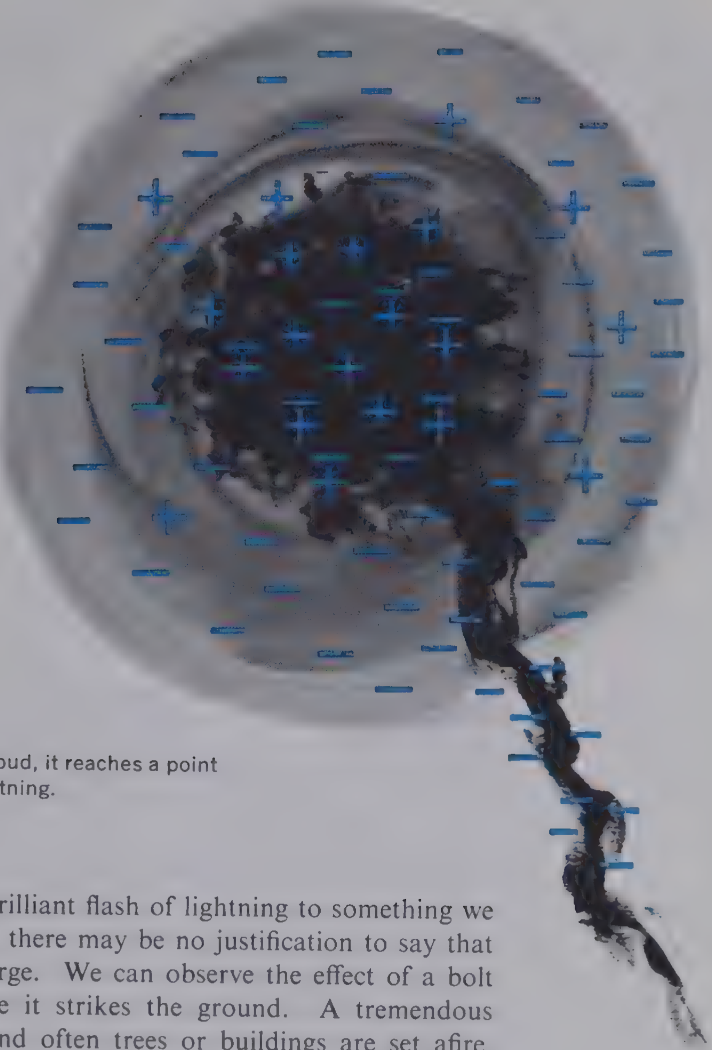
Energy radiating from the sun evaporates water and heats the air. This solar energy is also transformed into other kinds of energy. The tiny bits of water that evaporate rise and rejoin each other as they condense to make a cloud. The motion of the water particles produces a body of unbalanced charges in the cloud, somewhat the way rubbing a balloon produces an accumulation of charges. The air through which the water rises acts as an insulator to this accumulating charge. As the charge accumulates, a point is reached at which the air can no longer insulate against it. Individual particles of air, each containing equal numbers of opposite charges, are stressed by the forces of the electric field. When the stress becomes too great, charges of one kind, the electrons, are ripped from air particles and begin to stream through the surrounding air to neutralize the charge difference. Like a stone in the force field of the earth, charged particles tend to "fall"—some upward and some downward—toward the center of the opposite charge. They do not move far before striking other air particles, dislodging more charges, and eventually an avalanche of particles results—a bolt of lightning is formed.

The great energy reserve accumulated during the growth of the storm is now reduced. The detached charges again respond to charges on individual air particles and regain their original neutral arrangement. As the charges fall into place they become a source of radiant energy and release as light the energy gained when they were torn apart.

When electrons are removed from neutral atoms, as they are during a thunderstorm, the atoms are called *ions*. In fact, every charged particle of matter is known as an ion. Because they are charged, the ions are accelerated by an electric field. The very high speed they attain makes high-energy collisions possible. These collisions can result in a combination of particles to form new compounds. For example, nitrogen and oxygen gases normally exist together in the atmosphere without interaction. However, if their particles become charged as ions, they can unite to form new nitrogen compounds necessary for the nutrition of green plants. Rain dissolves these new compounds and brings them to the soil, making them available to plants. A form of oxygen called *ozone* is also a product of an electric storm and provides the odor we associate with electric sparks. Ozone is partially responsible for the fresh odor that accompanies rainfall.

Our explanation of a thunderstorm uses a particle motion theory for matter. In addition to bits of water, we talked about particles of air, tiny electrical charges called electrons, and somewhat larger charged particles from the air called ions. [1] Can we build an explanation based on a caloric fluid theory for the series of events in a thunderstorm? Can we explain lightning starting with radiation from the sun and the existence of the electric fields to the transport of electric charges, and finally to the emission of light?

Let's try to explain lightning using the caloric fluid theory. Radiation from the sun warms the air; that is, caloric fluid accumulates on the water and the air as the cloud develops. Then the caloric fluid is carried by the rising gases to the cloud level. As the divided water recombines to form the larger visible droplets that make up a cloud, an excess of caloric fluid accumulates. Finally, so much caloric fluid has gathered in the cloud that it cannot be contained. This is the bright discharge that we see connecting cloud to earth or cloud to cloud. Thus, we can offer an explanation of lightning using a caloric fluid theory. That we can do so is a warning that we need to examine critically every theory.



As electric charge builds up in the cloud, it reaches a point when it bursts from the cloud as lightning.

Unless we can relate the brilliant flash of lightning to something we already know about electricity, there may be no justification to say that it is, in fact, an electric discharge. We can observe the effect of a bolt of lightning on trees or where it strikes the ground. A tremendous quantity of heat is released, and often trees or buildings are set afire. Benjamin Franklin's experiments convinced him that lightning was the same sort of phenomenon as the electricity other scientists of his time had studied in their laboratories. [2] What other evidence can you think of from your own experience that relates lightning to electricity?

We have not explained the creation of new substances, such as the compounds made of nitrogen and oxygen. We believe that the chemical changes that produce these substances are the result of the electric activity of a thunderstorm. However, the events of a thunderstorm are so vast and so rapid that it is difficult for us to be sure. Let's attempt to bring chemical reactions down to a scale at which we can study them more carefully. We will attempt to refine our theories about energy exchanges as they relate to chemical energy.

Does the reaction in the metal of the nail produce any heat?

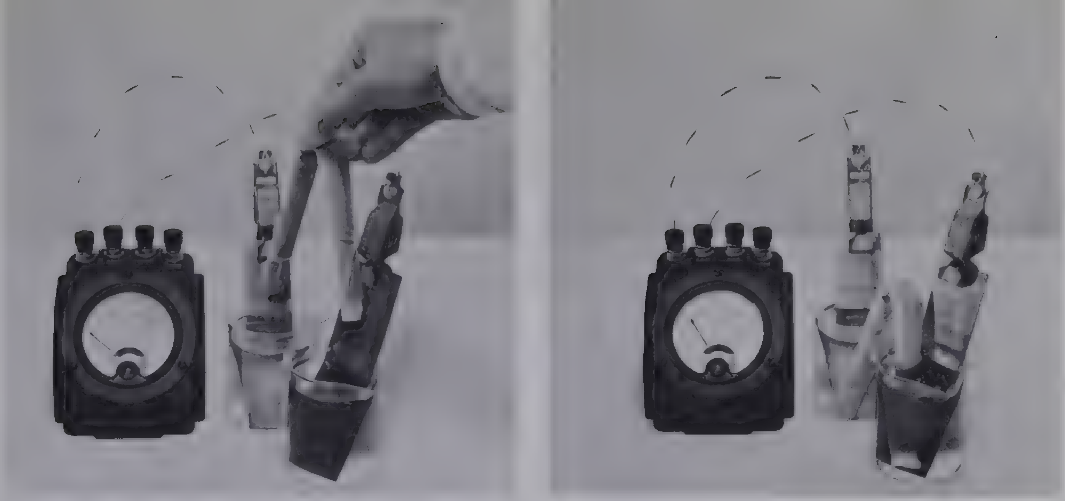
11–2 A CHEMICAL CELL

Recall Experiment 4–5 in which you put a nail in copper sulfate solution and found that it soon became coated with copper crystals. The blue color of the solution faded in the process. At that time you were interested primarily in the crystals, but now you can profit from a closer look at the chemical action involved.

11 1 To observe this chemical action you will need an iron nail, a test tube of copper sulfate solution, and a thermometer. Put the thermometer and the nail in the copper sulfate solution. Observe any changes in the nail, the solution, and the thermometer reading. An increase in the thermometer reading may indicate that energy is being released as heat.

Repeat the experiment using a piece of zinc instead of the nail. Record your observations of any changes that take place. It is not at all apparent that electric energy is being released in these experiments with iron and zinc.

11 2 To find out whether chemical energy can be converted to electrical energy, we need to assemble the same materials in a different way. You will need two 250-ml beakers, a 1- by 10-cm strip of copper, a 1- by 10-cm strip of zinc, two wires with alligator clips, a rolled paper towel, 200 ml of copper sulfate solution, a vacuum tube electroscope, 200 ml of zinc nitrate solution, and 50 ml of ammonium nitrate solution. Pour the copper sulfate solution into one beaker and add the copper strip. Attach the copper strip to one of the wires with the alligator clips. Add the zinc nitrate solution to the second beaker. Place the zinc strip into this solution



If the toweling were in only one solution, would the meter register a voltage?

and clip the other wire to the zinc. This setup is called a *cell.* Holding one wire in each hand, touch the tips of both wires to your tongue at the same time. If there is any electric energy generated by this cell under these circumstances, you will experience a metallic taste you would not notice if you used copper wire only. **(CAUTION:** This method for detecting the flow of current is *not* safe with many cells. Most commercial cells produce a dangerously large current. This method should be used only in cells like the one we have constructed here or in small dry cells. Using an electric meter that registers current is always a safer method.)

Wet the paper towel with the ammonium nitrate solution. Squeeze out the excess solution. Insert one end of this paper towel into each of the two beakers of solution and again test the two wires with your tongue for any evidence of electric energy. With the vacuum tube electroscope used in Chapter 10, try to determine whether the copper or the zinc strip is the positive terminal of your cell by connecting the wires to the cathode and the grid. (A voltmeter or galvanometer may be used instead of the electroscope.) If an effect is not immediately apparent, connect several cells in series (the wire from the copper strip of one pair to the wire from the zinc strip of another). This increases the available electric potential and, with enough cells combined, it should be possible to determine the polarity. When two or more cells are connected in this manner, the combination is known as a *battery.*

If your experiment was successful, you were able to establish that electric energy was released. Since some of the cell components were

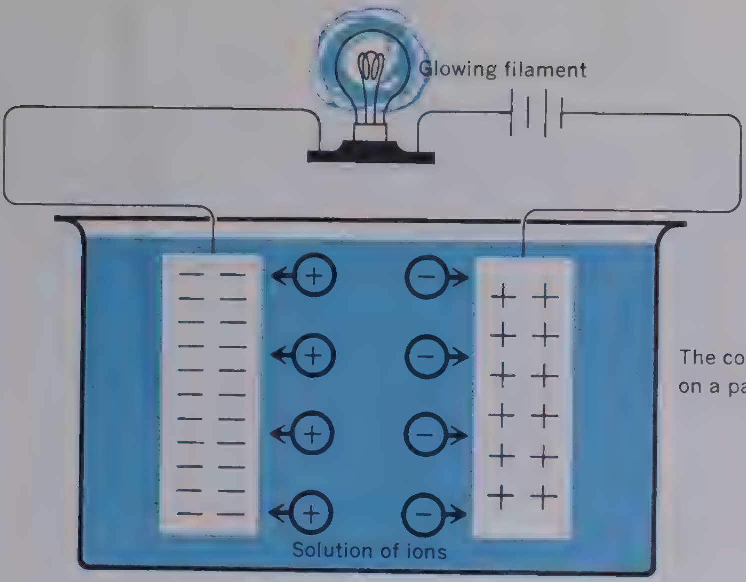
Is the solution being tested an electrolyte?

changed to new substances, we call this a *chemical reaction*, and the electric energy released or the heat noted earlier came from energy stored in the chemicals. We call this *chemical energy*, and during the period of change a chemical reaction is taking place.

Before we attempt to develop a theory for chemical reactions, we should take another experimental view of solutions.

11–3 THE CONDUCTIVITY OF SOLUTIONS

If electric current is to flow in any part of a circuit and allow us to utilize electric energy, there must be a closed path—the circuit must be "complete." Obviously, then, the solutions we used in the chemical cell must be a part of this path and must conduct electric current. We may explore the conductivity of solutions with a relatively simple apparatus. In the circuit connecting a flashlight lamp with a dry cell, we may introduce two electrodes which can be dipped into a liquid. If the solution is a good conductor it will complete the circuit and allow the bulb to light as shown in the illustration. To find out you will need the conductivity tester, ten 250-ml beakers, 100 ml of alcohol, tap water, distilled water, and 100-ml solutions of copper sulfate, zinc nitrate, ammonium nitrate, sodium chloride, sugar, acetic acid (concentrated), and acetic acid (dilute). Test each solution and liquid to see if it completes the circuit.



The conductivity of solutions based on a particle theory.

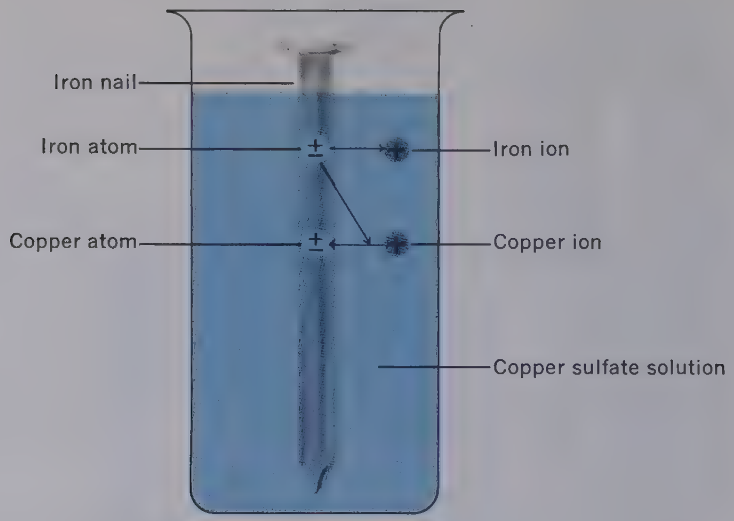
(Be sure to clean the electrodes thoroughly after each test and before placing them in the next liquid. This will prevent unnecessary contamination of the liquids which could affect the results.) If you connect the electrodes to each other by a wire, the bulb will glow brightly. If a solution completes the circuit, then the brilliance with which the bulb glows is a crude indication of the conductivity of the solution.

Liquids and solutions that conduct an electric current readily are called *electrolytes*; nonconducting liquids and solutions are known as *nonelectrolytes*. Almost all materials will conduct minute currents, but these may be too small to light the lamp. The electroscope could also be used to compare these liquids and solutions, as only the tiniest currents have to flow between the grid and the cathode of the vacuum tube to change the brightness of the neon bulb in the circuit.

$\boxed{3}$ As you examine the data collected from this experiment, can you see a trend that enables you to make a tentative statement about the kinds of liquids that are electrolytes and nonelectrolytes? $\boxed{4}$ What effect did the concentration of acetic acid have on its conducting ability?

Experiments with the chemical cell and with the conductivity of solutions supply evidence that chemical energy and electric energy are closely related in chemical reaction. The change of electric energy to chemical energy is not confined to thunderstorms, but is a part of all chemical reactions.

$\boxed{5}$ How can our theories of matter and energy be applied to chemical reactions and conductivity?



We can use the particle theory to build a description of the action at the surface of the nail.

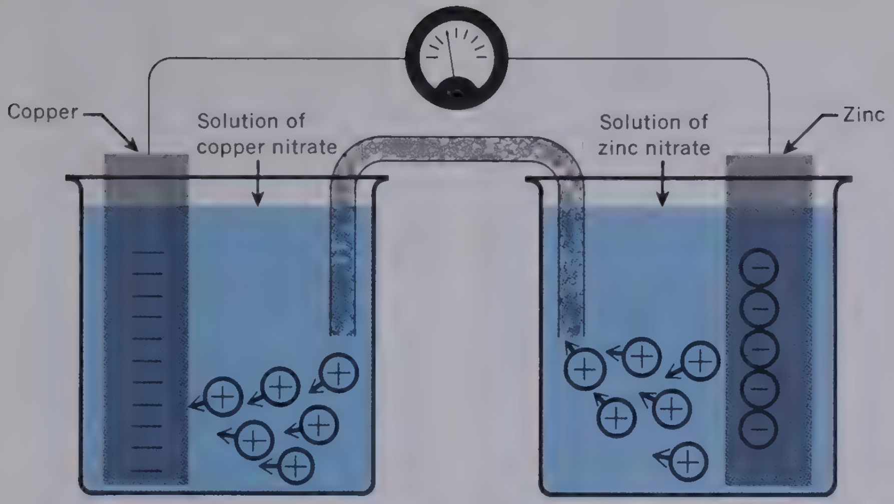
11–4 CHEMICAL CELLS, CONDUCTIVITY, AND IONS

In Section 11–1 we made reference to charged particles called ions. These particles are the key to an explanation of the relationship between chemical energy and electric energy in terms of a particle theory of matter.

Several solutes used in our study of solutions belong to a class of compounds chemists call *salts.* When salts are dissolved in water they are good conductors of electricity. Copper sulfate is a typical salt; its conductivity in water solution can be explained by saying that positive ions of copper and negatively charged sulfate ions are present.

The chemical change that occurred when the nail was placed in copper sulfate solution can be explained by ion formation. When the nail was placed in solution, a few particles of iron dissolved. In dissolving, iron atoms from the nail lost electrons to form iron ions. These electrons were transferred to the copper ions. When this happened, the copper ions became neutral copper atoms and were deposited on the surface of the nail as metal crystals. The iron ions replaced the copper ions to form a new compound, iron sulfate, which is colorless in solution.

Zinc reacts in much the same manner as iron. Electrons are transferred from zinc to copper with release of heat energy and electric energy. [6] Can we expect all chemical changes to be accompanied by an energy change? [7] Is energy absorbed in some reactions and released in others in the form of heat, light, or electricity?



In one beaker zinc atoms give electrons to the circuit as they become zinc ions. The electrons travel through the circuit to the other beaker where they are picked up by copper ions which become copper atoms.

In the chemical cell we separated the zinc plate from direct contact with the copper sulfate solution. A direct exchange of electrons between copper ions and neutral zinc atoms was not possible. What was possible, however, was the passage of electrons along the wire from the zinc plate to the copper plate that was in contact with the copper ions. When the circuit was completed through an outside wire and across the bridge, ions inside the solutions reacted just as they did when the zinc plate was in contact with the copper sulfate solution. One advantage of the chemical cell is that it provides detectable and usable electric energy. Dry cells used in flashlights and storage batteries used in automobiles are examples of practical applications of chemical cells as sources of electric energy.

There was another effect observed in the simple experiment in which the nail was placed in the copper sulfate solution. The nail became plated with copper, although the quality of the plating job was poor. In commercial copper plating, a source of electric energy that is separate from the reaction is ordinarily used. In our experiment the nail was a part of the cell and the source of energy that caused the reaction. The iron went through a smelting process to become metallic iron. Energy stored in it during smelting is now being used to connect the copper ions to atoms of copper metal.

The difference between the effect noted in the cell before and after inserting the connecting bridge must be related to the contribution the bridge makes to the total cell. 8 Would the same function have been performed by a piece of copper wire or by a cotton roll soaked in a solution of any other salt?

Chemical change

Physical change

11–5 CHEMICAL AND PHYSICAL CHANGES COMPARED

The action of cells is an excellent example of the nature of chemical reactions involving ions. One characteristic of all chemical reactions is that the products formed are quite different from the original substances. The products are not easily converted back to the original or reactant forms. By contrast, the boiling of water, where the product, steam, does not resemble the liquid from which it was produced, can be reversed quite readily by removal of the heat energy that is in the steam. Changes of this kind are generally called *physical changes*, or *phase changes*. When iron or zinc is placed in copper sulfate solution it is much more difficult to reverse the reaction.

In chemical changes, the particles of several substances can combine, or complex particles can divide to form new substances. The particles are no longer the same and the products show properties that may be quite different from the properties of the original substances. A striking example of such a chemical change occurred when you introduced ammonia and hydrogen chloride into a glass tube. Here the reacting substances were both colorless gases. When particles of the gases met in the tube, the solid white substance, ammonium chloride, was formed. Like the products in the electric cell, ammonium chloride is not easily converted to the original gases of which it is composed. [A, B]

Evaporating the water after the reaction stops.

11–6 ENERGY IN CHEMICAL REACTIONS

Energy is an important factor in chemical changes, just as it is in physical changes. The amount of energy involved in any reaction is one factor that determines whether the reaction is classified as a physical or a chemical change. Most chemical changes involve a great deal more energy than physical changes. With zinc in copper sulfate solution we noted an increase in the temperature as the reaction continued. Heat energy must be present when such a change in temperature occurs. Since no other source of heat energy is present, we could conclude that the reaction itself is the source of the heat energy. Energy was apparently stored in the original substances. This might be one of the reasons that the reaction took place; the lower energy state may be a more natural condition. There are, by contrast, many reactions in which a lower energy condition does not seem to be the governing factor.

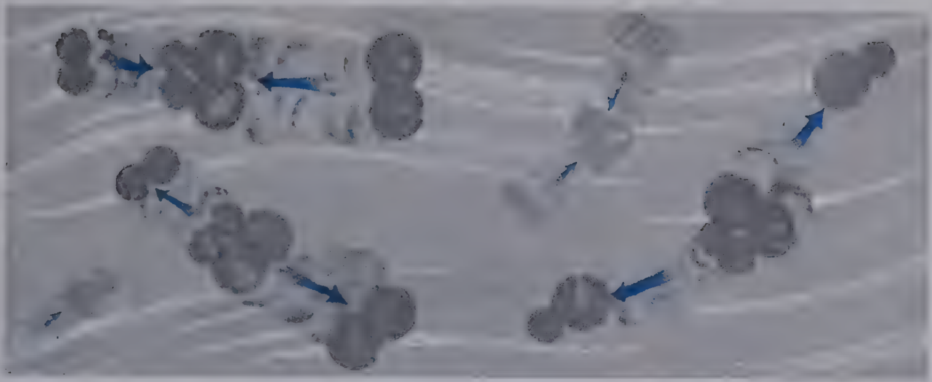
We can easily observe one of these reactions. You will need a beaker, Alka-Seltzer tablets, a mortar and pestle, and a thermometer. Put 100 ml of water in a beaker and determine its temperature. Then add an Alka-Seltzer tablet to the water. Observe what happens as the tablet dissolves, and note the temperature when all changes seem to have stopped. Over a low heat, evaporate the water from this mixture and examine what is left. [9] Does it look like the original tablet you added to the water? [10] Is the solid remaining the only product of this reaction? 11 4

Grind up another Alka-Seltzer tablet in a mortar and compare the resulting powder with dry residue from the reaction with water. Add a few drops of water to some of the powdered tablet and compare the effect on it with that from adding water to a sample of the dry residue. [11] Going by our assumption that in a chemical reaction the products are different from the original substances, in which of these cases does a chemical reaction occur? The change of temperature is not what we might have expected from the results we observed in the cell or when metal was added to the acid. [12] Can you suggest any explanation for the difference? One clue may be found if you recall what happened when you dissolved hypo in water in the study of solutions.

Let's return again to a consideration of the reaction when ammonia and hydrogen chloride combined. We observed that the rate of this reaction seemed to depend on the temperature of the air through which these gases moved. When the materials were warm, the gases traveled down the tube faster and reacted sooner than they did when the tube and gases were cold. This might be explained by using the theory for gases that we have developed. The gas particles are in rapid random motion. At higher temperatures the motion is more rapid than at lower temperatures. The particles have more kinetic energy, and as they move about they collide. Both the frequency and the force of collision are greater at higher temperatures.

High-energy collisions between particles of ammonia and hydrogen chloride resulted in the formation of a single new substance, ammonium chloride. From this we can conclude that the colliding particles stayed together rather than bouncing apart. However, many collisions lead to two or more new particles. At higher temperatures high-energy collisions are more likely to occur than at low temperatures. Thus, there is a greater probability of collisions of sufficiently high energy to disturb the organization of particles than we would find at lower temperatures. Reorganization into new kinds of particles becomes possible. In many reactions, the collisions disturb the internal structure of the colliding particles so that a number of new kinds of particles are formed instead of two particles becoming one, as they do in the ammonia-hydrogen chloride reaction.

In looking back at other situations we have encountered, we find examples of the energy of the particles at room temperature not being great enough to cause chemical reaction between the substances, even though they were well mixed. Alcohol vapor or the combustible gas in the burner of a stove do not combine chemically with air until they are heated by the flame of a match or the spark of a lighter. Then, in the chemical change we call burning, enough energy is released so that



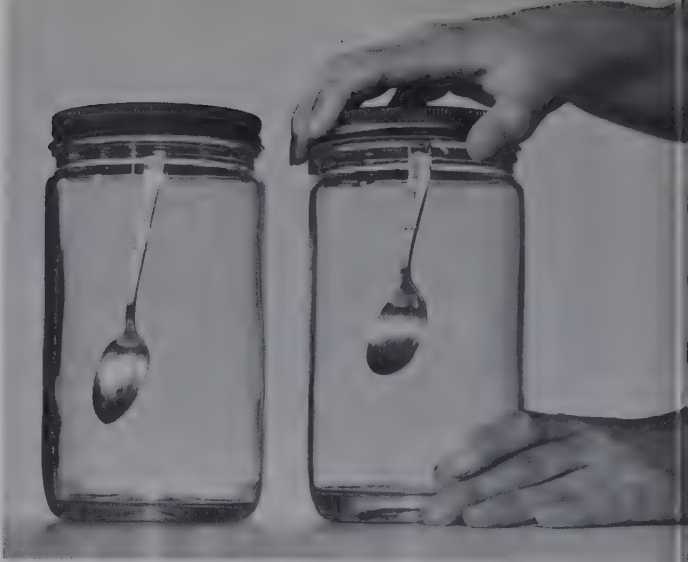
When particles collide they may bounce apart, react to form a single new particle, or react to form two new kinds of particles.

additional particles can react and the process becomes continuous. We describe the lowest temperature at which burning can occur as the *kindling temperature*.

In burning fuels, stored energy is released in a continuing reaction to warm our homes or cook our food. All common fuels are products of living things; organisms may have stored this energy during the growing process last summer, or a few million years ago. They are useful fuels mainly because the stored energy can be released from them under controlled conditions that permit the energy to be applied to our needs. The sun is the original source of the energy and its radiation is bound into the chemical combinations we call carbohydrates, fats, and proteins. The chain of events that produces these complex combinations of the elements carbon, oxygen, hydrogen, and nitrogen is called photosynthesis. When fuels are used, their chemical energy is converted to heat energy.

There are a number of important reactions that will not continue unless energy in some form is continuously supplied to the reaction. The refining of natural ores to produce useful metals requires the addition of energy. To obtain the metals from their native ores, we must supply energy that was released when the mineral ores were formed. All the processes of the steel industry start with the separation of iron from its ore.

By contrast, many reactions are spontaneous. From our experience with the rusting of iron, we know that water has a distinct effect on the corrosion of metals. The presence of other substances in the air, such as salt or sulfur compounds, also speeds the corrosion process.

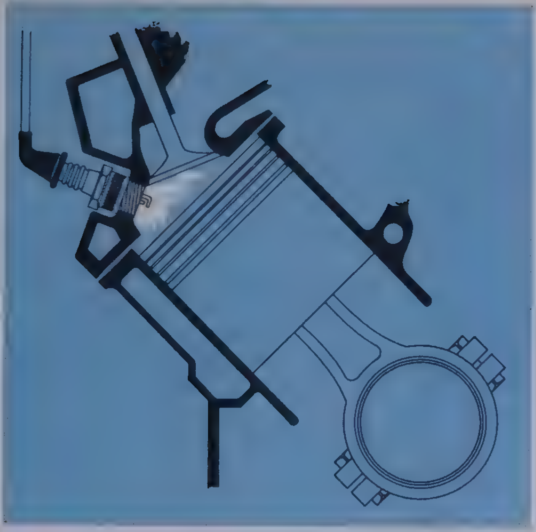


Two simple controlled experiments can be done to show the influence of water and hydrogen sulfide on corrosion. For these experiments you will need four quart-sized jars with tight-fitting lids, two balls of steel wool, two pieces of silver-plated flatware, about 10 ml of concentrated sulfuric acid, 5 g of ferrous sulfide, 50 ml of water, and several pieces of string.

Put 25 ml of water in one jar. Hang one ball of steel wool just below the lip and put the lid on tightly. Put about 5 ml of sulfuric acid in a second jar, without having the acid touch the upper sides of the jar. Hang a second ball of steel wool in this jar and put the lid on tightly. In the third jar put about 25 ml of water, then add 2 ml of sulfuric acid. Be sure that the walls of the jar are dry, then carefully add several pieces of ferrous sulfide. Hang one piece of flatware above this mixture and put the lid on the jar. For this jar, leave the lid loose since some gas will form in the solution. In the fourth jar, put about 5 ml of sulfuric acid and hang the second piece of flatware above the acid. Put the lid on tightly. Put all four jars in a place where they will be safe and undisturbed and leave them for several days. Make observations from time to time to see if anything is happening to the steel wool and the flatware. [C, D, E]

11–7 CHEMICAL REACTIONS OF NONIONIC SUBSTANCES

It is interesting to note that the same elements, iron and oxygen, combine during the slow process of rusting and in an intensely hot and rapid process



The power of an internal combustion engine comes from controlled explosions.

when finely divided iron is heated in the presence of abundant oxygen. The bright sparks that fall from a grinding wheel when a steel tool is sharpened are hot enough to react rapidly with air to form an iron oxide. The variety of conditions under which iron and oxygen react to form iron oxides leads us to believe that the reaction is affected by the conditions that surround it, such as temperature.

It would seem reasonable to assume that rusting and other corrosion processes are reactions involving ions, but the burning of a fuel, such as gasoline or alcohol, may be a reaction that does not involve ions. Remember that conductivity measurements showed that particles of some liquids are not charged and cannot be used to carry an electric current. Solutions of alcohol and sugar in water are examples. However, these two substances do burn and the products are sufficiently different from the reactants to give us reason to believe that burning is a chemical reaction. The process of burning of such fuel materials is not an ionic reaction. The reorganization of the particles of burning substances may involve the interaction of electrically charged parts, but it is not as easy to investigate and establish the nature of such reactions as it is with ionic reactions.

It is easier to study the effects of concentration and temperature on the rate of reaction of ions than on nonionic substances. We know how rapidly some fuels will burn if the correct conditions of temperature and concentration are obtained. Such rapid burning is usually called an *explosion.*

How does the temperature affect the reaction rate?

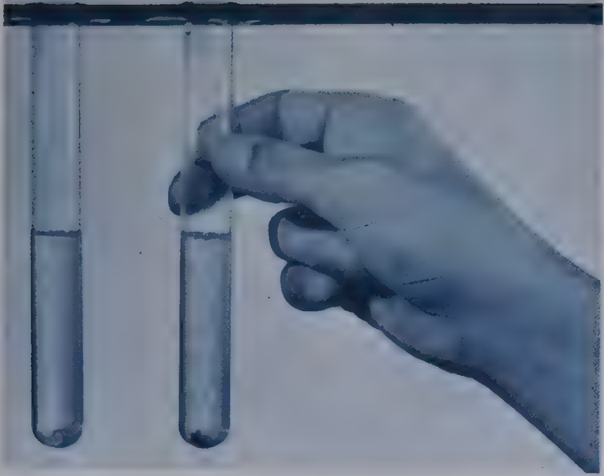
11–8 REACTION RATE AND TEMPERATURE

We have suggested that temperature has an effect on the rate of a chemical reaction. Not only is it necessary that the two substances reacting—the reactants—be brought to some initial temperature (level of heat intensity) before burning can occur, but heat in the surroundings may influence the rate at which the changes take place.

If we could burn iron at very high temperatures and compare the rate with that of rusting of iron at room temperature, the difference would be startling. But we cannot do this conveniently. We may, however, obtain useful data about temperature effects on reaction rate with the Alka-Seltzer reaction used earlier. A graph of the results may give some indication of whether a relationship exists between temperature and the time for a given reaction to take place. Each student in the class should determine the time required for a tablet to dissolve completely in water at a particular temperature somewhere between the freezing and boiling points of water. If each student tests the reaction time at a temperature different from others in the group, enough data may be obtained to construct a graph.

11
6

For the experiment you will need one 250-ml beaker, a thermometer, 100 ml of water, and one Alka-Seltzer tablet. Put the measured 100 ml of water in the beaker; record the temperature. Drop the tablet into the beaker and note, to the nearest second, when the last solid portion of the tablet seems to have finished reacting with the water. There may be some undissolved material which will fall to the bottom of the beaker, while other material rises to the surface and escapes as gas. [13] Why does some material escape while some does not dissolve? Use the observations of your classmates to plot a graph of the time of reaction as a function of



How does concentration affect the reaction rate?

temperature. Actually this graph does not describe the *rate* of reaction directly; instead it gives the time to complete the reaction. The shorter the time, the faster the rate. To obtain values related to the rate, divide the time for each temperature into the time taken for the slowest reaction. Plot these values as a function of temperature. [14] For this study, should the initial temperature or the final temperature of the liquid be used? [15] Would an average temperature be better? [16] What effect does temperature have on reaction rate as indicated by this experiment?

In the above experiments everyone used the same quantity of water and Alka-Seltzer. The only variable was the temperature. It is possible that the quantity or concentration of a reactant may also influence reaction rate. [17] Would a second tablet dissolve in the solution as rapidly as the first one if the temperature were the same at the beginning? [18] Based on these observations, how long do you think it would take for two tablets to dissolve if they were put in the water simultaneously? [F]

11–9 REACTION RATE AND CONCENTRATION

We may do another experiment to study the effects of concentration using magnesium ribbon and acid. You will need two small test tubes, eight strips of magnesium ribbon 2 cm in length, a dropper, two thermometers, a test tube rack, a graduated cylinder, dilute hydrochloric acid, and water. To one test tube add ten drops of dilute hydrochloric acid and to the other tube add 20 drops. Then add enough water to make 10 ml of solution in each tube. Shake both tubes to mix the acid and water, then drop a 2-cm strip of magnesium ribbon into each. Observe the reaction rate in each tube and record the results. Also note and record the initial 11 7

Can a solution react with an unlimited number of pieces of magnesium?

and final temperatures of the solutions. 19 If you had rolled the strip of magnesium ribbon into a tight spiral, do you think the time required to dissolve would have been different?

Again dissolve a second similar straight piece of magnesium ribbon in the test tube that contains the more dilute acid solution and note the reaction time. Repeat this process with a third strip in the more dilute acid solution and plot the rates as a function of the number of strips dissolved. 20 Can you predict the total number of strips that could be dissolved by 10 drops of the dilute acid by determining the point at which the dissolving rate would seem to be zero? Such an extrapolation suggests that each unit of magnesium ribbon "uses up" a certain amount of the available agent in the acid. The implications of this are tremendous. It suggests that there is a definite relationship between reacting quantities of the metal and the acid. In terms of our particle theory, it would indicate that, as more and more metal was consumed, there were fewer available particles of acid remaining to act on the ribbon. The chances of collision of remaining acid particles with metal particles becomes less and less. Eventually the number of collisions becomes so low that reaction is no longer observed.

11–10 REACTION RATE AND SURFACE AREA

In question 19 above we implied that the amount of surface area may be a factor in the reaction rate. Recall the effect surface area had on the dissolving rate of salts studied in our consideration of solutions in Chapter 3.

If chemical reactions result from collisions, then it would seem that the rate of a reaction would be greater, all else being the same, if the surface area in contact were also greater.

Break an Alka-Seltzer tablet into eight or more pieces. Determine the time required for the broken tablet to react completely. Compare this with the dissolving time required for an unbroken tablet and one that has been crushed to powder. [21] What have you done to the surface area of a tablet by breaking it up? [22] Does the experiment you have just completed support the particle theory for the structure of matter? 11 8

11–11 REACTION RATE AND CATALYSTS

One other significant factor influences the rate at which many reactions occur. A substance that alters the rate of a reaction and that is not permanently changed by the reaction is called a *catalyst*.

To learn more about catalysts, you will need six test tubes, 30 ml of 3% hydrogen peroxide solution, and tiny quantities of manganese dioxide, copper metal, zinc metal, sodium chloride, and calcium carbonate. Divide the hydrogen peroxide solution among the six test tubes. [23] Is there any evidence of separation of the hydrogen peroxide into water and oxygen gas? Add one of the other substances to each test tube using the sixth as a control. Observe carefully for evidence of gas formation. [24] Which substances are acting as catalysts? [25] Are some of the substances more effective catalysts than others? 11 9

Catalysts that are a part of the chemical processes of life are called *enzymes*. Most reactions that go on within living systems are possible only because of these substances. If enzymes were not present, the reactions that they catalyze would go so slowly that living organisms could not exist.

We must realize that catalysts do not affect the *proportions* of reactants and products that exist in a chemical system. They affect only the *rate* of the reaction and the time it will take the reaction to reach equilibrium.

Both enzymes and catalysts can be made ineffective by poisons that destroy the catalytic capacity. Many such poisons react with enzymes in the human system and so seriously interfere with the normal operation of the system that death results. This is a very good reason why substances of unknown composition should never be put in your mouth.

11–12 REACTION RATE AND ELECTRICAL ENERGY

We have observed the effect of temperature on reaction rate. The effect of temperature is actually the effect of kinetic energy. Will other forms of energy produce similar results? Can electrical energy from a battery alter the rate of chemical reaction?

11 10 To answer these questions we will study the rate at which electrical energy decomposes water. This is a process called *electrolysis*. You will need two test tubes, a 250-ml beaker, conductivity testers, electrodes, and a dry cell. Add dilute sulfuric acid to 200 ml of water to make it a good conductor. Use the conductivity tester you used in Section 11–3 to check the solution. Fill the two test tubes with this solution and invert them over the electrodes, as shown in the illustration. Connect the two electrodes to the opposite poles of a battery or other source of direct current. As the reaction proceeds, compare the quantities of gas that collect in the tubes. Alter the amount of current and observe reaction rates. Make a note of the amounts of gas collected at each electrode.

The gases that collect in the two tubes are very different, although they both appear as colorless substances that show their presence only by displacing some of the liquid from the tubes. One gas is hydrogen, which has the lowest density of any known substance. The other gas is oxygen. Hydrogen may be kept with very little mixing in an inverted open test tube; air, since it is heavy in comparison, will not mix with it. Instead, the hydrogen tends to float on the air. If ignited in an inverted tube, hydrogen gas will burn calmly at the mouth of the tube. The bluish

What do you notice about the volumes of the two gases?

flame is nearly invisible. If hydrogen gas is mixed with the oxygen in the other tube, the mixture explodes when ignited by a flame or spark.

When hydrogen is burned in oxygen, water vapor is the only product. The chemical process initiated by the electric current is reversed. The sound and the temperature change that accompany the reaction are indications that stored energy is released. Recently it has become possible to use this combining reaction to produce electrical energy in a device called a fuel cell. This provides a very convenient source of electricity for space vehicles.

11–13 CHEMICAL EQUILIBRIUM

Our observations of chemical reactions have been focused on those that go in one direction. In the physical changes we studied earlier, however, it seemed that they all reached states of equilibrium, but that the changes could be reversed. Does this provide us with another distinction between chemical and physical processes? It turns out that the answer is no. If we extended our investigations we would find that chemical reactions also lead to equilibrium states. Equilibrium is usually more difficult to detect and observe in relation to chemical reactions. To know that a chemical reaction is taking place, we have to observe some change in the materials involved. In Chapter 4 we produced crystals of several salts on a microscope slide by using chemical action. The reaction proceeded toward the formation of the crystals because the salts (barium sulfate, magnesium sulfate, and strontium sulfate) were not sufficiently soluble in water for much to remain in solution. A great amount of each substance precipitated as a solid. The precipitate was evidence of chemical reaction. In later experiments, such as those with the Alka-Seltzer, a gas was released. Equilibrium between the reactants and the products could not occur because of the escape of one of the products.

In other experiments you have done, such as in the chemical cell or the replacement reactions of zinc and copper in the blue copper sulfate solution, energy was lost. In the latter reaction it seemed to be heat energy, but this was not the original energy released. Instead, electric energy was converted to heat. $\boxed{26}$ Would warming the substance return the copper to solution? $\boxed{27}$ Could an electric current be used to reverse the reaction?

When a chemical cell is not in use, an equilibrium does exist. Copper ions become atoms and atoms become ions at exactly the same rate, and the zinc also changes from atom to ion and back again. The total number of ions and of atoms of each kind is constant. When the circuit is completed, energy is removed and the reaction produces many more zinc ions and copper atoms. If we could, with a higher voltage, cause the electric current to flow in the opposite direction, we could replace some of the energy that escapes and, to some extent, reverse the reaction.

The dry cell used in flashlights or electronic equipment is an equilibrium system until the circuit is completed and some energy is removed. The lead-acid cells in the storage batteries in cars are in equilibrium during inactive periods. By adding electric energy, the direction of chemical action goes one way. The battery is charged. As electric energy is taken from the storage battery, the chemical action proceeds in reverse to the charging direction.

In reactions that produce a gas, the reaction can sometimes be brought to a state of equilibrium by preventing the escape of the gas evolved.

11 11 For a simple experiment you will need two small cylindrical containers with tight-fitting caps, transparent adhesive tape, cool water, a timer, and two Alka-Seltzer tablets. For safety, spirally wrap one of the bottles with the tape from end to end, overlapping the tape about half of its width each turn. This double layer of tape will give the bottle great strength against splitting. Fill both bottles with tap water, being sure that the water in each bottle is the same temperature and that the bottles

ng a bottle for Experiment 11–11.

are filled to the brim. Drop one Alka-Seltzer tablet in each bottle simultaneously and quickly screw the cap tightly on the taped bottle; leave the second bottle open. Record the time for the reaction to come to completion in each bottle. When the reaction seems to have stopped in the closed bottle, loosen the cap slightly to allow some gas to escape, then tighten it securely again. Observe the results of discharging this bit of gas. When the reaction appears to have stopped again, release another bit of the compressed gas. Continue this procedure until no reaction is evident when the pressure is released. $\boxed{28}$ Was there any indication during this experiment that the gaseous product may have been in equilibrium with the dissolved gas? $\boxed{29}$ Is there any indication that the chemical reaction that released the gas might have been arrested by this increased pressure? $\boxed{30}$ Do you think pressure might affect the rate of reaction in a system where gas was not produced?

If a gas is produced in a chemical reaction, there is no opportunity for equilibrium to occur unless the gas is kept in the reacting system. When the pressure is released by the escape of gas, more gas is formed in an effort to restore the pressure and thus the equilibrium. This is what happened in the sealed Alka-Seltzer bottle. As you allowed gas to escape, more formed to restore equilibrium. A French chemist, Henri Louis Le Chatelier, noticed this tendency toward equilibrium in reactions and stated the basic idea of chemical equilibrium in what is now known as

Le Chatelier's Principle: *When any system in equilibrium is disturbed, the system acts in such a way as to counteract the change and restore an equilibrium condition.*

In precipitation reactions, equilibrium does not occur between the dissolved and solid substances until a considerable portion of the solid has formed. To cause some or all of the solid to redissolve, some ions that are in solution must be removed, or their form must be changed. A good example of such a change involves copper hydroxide.

To experiment, you will need solutions of copper nitrate, sodium hydroxide, ammonia, a test tube, and several droppers. Put a few milliliters of the copper nitrate solution in the test tube. Add a small amount of sodium hydroxide solution drop by drop, but do not add more than five drops. Now add ammonia solution drop by drop. Describe what happened as you added each solution to the copper nitrate. [G, H, I]

11–14 WHY DO CHEMICAL REACTIONS OCCUR?

We have built a theory for matter that pictures it as being composed of very small particles. We call the smallest particles atoms. Sometimes atoms occur in groups we call molecules. During your experiments you have made many observations that indicated that these particles are in constant motion. They follow no predictable path, but move in a haphazard, random fashion like a butterfly darting about.

We can explain chemical reactions, then, by assuming that collisions occur between molecules and ions. The force of the collisions is the primary determiner of whether molecules react. Molecular collisions can be compared to automobile collisions. If two cars collide head on at high speeds, you know that damage (change) is far more likely to result than if they are moving slowly. This comparison helps us understand why increasing the temperature of a system, which speeds up the molecules, increases the chances that a reaction will occur.

Another factor that helps to determine whether a chemical reaction will occur is the number of particles present. For example, collisions are far more likely to occur on an expressway during rush hours than when the traffic is light. A third factor that helps determine whether a chemical reaction will occur is the way in which the particles collide. If two cars meet head on, damage (change) is much more extensive than it is if the collision is at right angles or is a sideswipe. We can conclude from this analogy that the most favorable conditions for a chemical reaction exist when many particles of high energy are crowded (concentrated) together.

In considering particles of the size that engage in chemical change, collision means that the force fields surrounding the particles interact. There is no evidence that the particles actually physically touch each other. The strongly repelling fields of like particles prevent a close approach at any energies (temperatures) we have experienced. The fields of oppositely charged particles tend to bring them together but they do not seem to actually make contact. This attraction is enough to hold them in association and leads to the shape and size of the particles. [J]

TAKING INVENTORY

Chemical energy has been added to the list of forms of energy that we have examined. Investigation of evidence for this kind of energy, its conversion into other forms of energy, and the effect of energy on chemical reactions has been made in the form of our study of the chemical cell, the conductivity of solutions, the comparing of chemical and physical changes, nonionic reactions, and the effect of various conditions upon the rate of reaction.

The particle motion theory and the caloric fluid theory have both been re-examined in an effort to judge whether one offers a better explanation for chemical energy and chemical reaction than the other. Although the caloric fluid theory can be used to develop an explanation for sparks and lightning, the particle motion theory offers a better basis for a consistent explanation of chemical energy and chemical reaction.

FURTHER EXPLORATIONS

A. One of the most interesting and widely used sets of chemical reactions is based on the response of silver compounds to light. Obtain small quantities of silver nitrate and sodium chloride solutions. Pour a few milliliters of each together in a test tube. Keep the resulting white substance, silver chloride, away from bright light. Pour it onto a piece of paper toweling or filter paper and spread it out. Let the paper dry, then with your hand or some other opaque object partially covering it, expose the rest of it to sunlight or some other strong light source. Is there evidence that the light has caused a chemical change?

If you want to investigate further, obtain a small amount of the solution known as "developer;" moisten the paper with it and note what happens.

B. What evidence did you observe in Experiment 11–4 that indicates that a chemical reaction occurred rather than phase changes only? Were phase changes also involved?

C. The reaction of ammonia and hydrogen chloride is a chemical reaction. Did you notice any evidence of energy involved in the reaction as you performed it as part of your study of Chapter 6? Try to design an experiment that will measure the heat associated with this reaction.

D. Apply the particle motion theory to describe what is happening to the particles during the experiments with Alka-Seltzer tablets. This will be a test of the value of the theory. Can you build a similar description based on the caloric fluid theory?

E. Test a solution of Alka-Seltzer in water for electric conductivity. If the solution is conductive, how might you go about identifying what ions are present?

F. Why is Alka-Seltzer in tablet form a good material to use in Experiment 11–4 on the effect of temperature on reaction rate? What would you have to do if the same substance were packed by the box in powdered form?

G. Concentration and gas pressure have been discussed as having an effect on the rate of chemical reactions. Are they really two different factors?

H. In many places in the book you have encountered the statement that heat is always transferred from hot objects to colder objects. Try to apply the idea of equilibrium to this statement. As you do so, keep in mind that heat is particle motion.

I. We have made many references to equilibrium. When you worked with ammonia and hydrogen chloride to form ammonium chloride, did you observe any evidence of a reaction opposed to equilibrium? Obtain a small amount of ammonium chloride, place it in a 20-by-150-mm test tube. Heat the test tube gently at first, then more strongly concentrating the heat on the place where the substance is lodged. Keep the upper end of the tube cool. What happens? Is there any indication that you have reversed the reaction? Can you be sure that you have caused a chemical reaction? Put another small amount of the ammonium chloride in an evaporating dish; weigh the dish and put it in an open area in the classroom. Leave it there for a week.

J. Use the particle theory to build a description of the way chemical reactions occur. Take into account that temperature and concentration affect the rate of reaction.

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12

Matter and Energy: What Does the Pattern Look Like?

12–1 The Ways of Science
12–2 A Review and Refinement of Our Concepts
12–3 The Kinetic-Molecular Theory: Our Best Explanation
12–4 Natural Tendencies in the Universe
12–5 The Sun: Our Source of Energy

12–1 THE WAYS OF SCIENCE

Have you ever been mountain climbing and noticed the peak ahead? The peak seems only a short distance away because the surface usually slopes away and you cannot see the crest. As you climb higher, you see more of the land below. If you pause along the way you are rewarded by a broader outlook at each successively higher level.

The journey through this book resembles a mountain climb. We have explored the paths of science with our minds. We have learned by examining selected phenomena at close range. But like the mountain climber, we should pause to look at what we have studied in its entirety. The pattern of science is a composite of numerous lesser patterns which fit together, each complementing the other.

You have become a more careful observer. Little things that might have escaped your attention under other circumstances have now become more important—a small change in volume, a slight warming, a new color, a different shape, a change in motion or an altered force, a compound formed, or a substance released. If we are alert to less obvious facets of our environment and curious about their place in a large, overall pattern, then our study of science is far more meaningful. We have sought to learn by careful experimentation the place of each individual event in a large pattern. We have made reasoned guesses and have asked significant questions to test our reasoning.

A person who notices the inconspicuous occurrences, wonders about their place in a total pattern of experience, and endeavors to see new relationships is a potential scientist. The journey through life is more exciting and rewarding for those who work in this manner, whether they plan to become scientists or not.

An x-ray diffraction photo of a ruby crystal which provides indirect evidence of regular arrangement of particles in crystals.

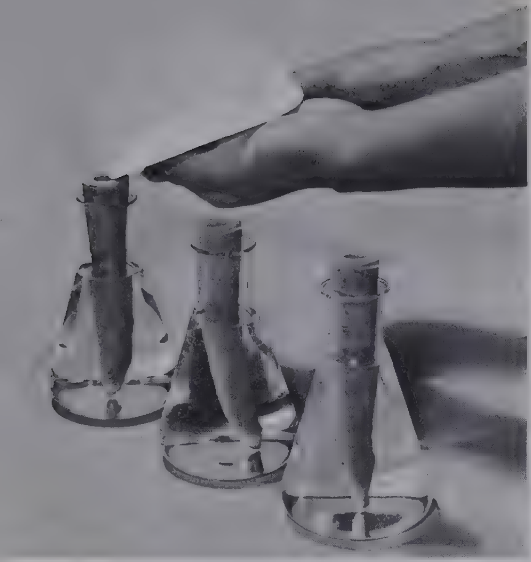
12–2 A REVIEW AND REFINEMENT OF OUR CONCEPTS

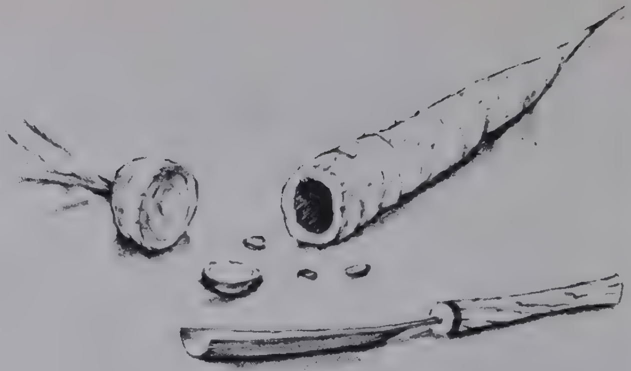
Let's review the significant outcomes of our past experiments and try to formulate the best possible statements about the structure of matter and the nature of energy.

Matter is Composed of Particles

Beginning with our early work with sulfur, its three forms, and the different densities of these forms, and continuing through the crystal models and the crystals themselves into our study of gases, electric charges, electric fields, and finally, into chemistry, the evidence is overwhelming that matter is composed of particles. In thinking of matter as being continuous, we say that matter can be divided indefinitely without changing the properties of the pieces; whereas a particle structure requires that there be a definite limit possible in the division process, that ultimately we come to a piece of matter which, if it were further divided, would be different in all of its properties. This does not mean that particles cannot be so related that they *appear* to have continuous characteristics.

The particles of matter are evidently extremely small, so small that it is impossible to observe them individually. Generally speaking, the particles to which we refer are called molecules. From our knowledge of particles we can extract a definition for the term *molecule*: the smallest particle of matter that has the properties of larger amounts of that substance. The uniformity of the characteristics and behavior of each individual substance leads us to believe that the molecules in any given kind of matter are more alike than grains of wheat or table-tennis balls. We cannot, at this point, make statements about the size, shape, or specific activities of these molecules. We would need experiments of an entirely





different nature than we have done if we were to be more specific about any of these characteristics. Remember that our evidence that molecules exist at all is indirect. Consider how much more difficult it would be to obtain more specific information about size and shape by indirect information. When we consider the size of molecules, do we refer to mass or to a linear measurement? And what is the meaning of shape when we refer to things that are below the level of our ability to see, even with a microscope?

It is possible to learn more about the relative sizes of molecules and about the mutual attraction of certain kinds of molecules if we perform experiments involving a process called *osmosis*. Osmosis is the passage of molecules through a barrier separating solutions of different concentrations. It is a form of diffusion, which, you will recall, is the process of intermingling between particles of different kinds of substances.

For this experiment you will need three carrots about the same size, three containers to hold the carrots, sugar, a sugar solution, and water. Hollow out the leaf end (top) of each carrot. Lower the first carrot into a container of water so that the top is above the water level. Then put a teaspoon of sugar in the depression in the carrot. Leave the depression empty in the second carrot and place it in another container filled with water at the same level as in the first container. Fill the depression in the third carrot with water and place it in the third container holding sugar solution. Over a period of several days record observations on what happens to the sugar and the water. Taste the solution in the container with the carrot in which you put sugar.

Predict the results expected if salt were substituted for the sugar. Set up an experiment to check your hypothesis. Test other materials such as cork, wood, eggshell, and parchment paper for evidence of osmosis. Examine these materials with a microscope and relate the observed behavior to their structure.

In plants, water passes from the soil into the roots at a much greater rate than the plant fluids pass out of the roots in the opposite direction. Try to devise an experiment to measure the rate at which fluids are transferred through a plant.

Fluid transfer occurs in all living systems. The tiny units of living tissue called cells carry on an extensive exchange of liquids and gases back and forth through their walls. This provides a supply of food nutrients and oxygen which enable the cells to live and also serves to remove waste products. The intestinal tissue of higher animals permits the passage of dissolved food into the digestive tract. Exchange of carbon dioxide and oxygen is carried on in lung tissues. All these changes require the expenditure of energy, and also give evidence that the particles must be quite tiny if they are able to pass through cell walls.

Another clue to particle size is obtained if we dilute a colored solution. Drop a tiny crystal of potassium permanganate into 50 ml of water and stir until the color is uniform. Pour 25 ml of this solution into another beaker and dilute it with 25 ml of water. Take half of this solution and dilute it to 50 ml with water. Continue this dilution process for as long as you can detect color. Compare the intensity of color in the different beakers. If the particles of potassium permanganate were large (as was 12 2 the original crystal) the color would appear in blobs. Since successive dilutions result in producing lighter and lighter color that is uniform through the solution, the potassium permanganate must have a tremendous ability to spread and scatter. This can be explained only if we assume the particles to be extremely small.

Molecules Are in Constant Motion

Evidence from studies of diffusion in liquids and gases indicate that for these changes to occur, molecules must be in constant, random motion. Potassium permanganate crystals in a container of water eventually become uniformly distributed throughout the water even though no stirring occurred. Ammonia gas and hydrogen chloride gas moved through a glass tube until their molecules collided with sufficient energy to form a new kind of molecule, ammonium chloride. You will recall that this reaction occurred in less time when the experiment was performed at high temperatures than at low ones.

Since the rate of motion of the molecules depends on their kinetic energy, and since heat increases the kinetic energy, it becomes obvious that the temperature of the system affects the rate of molecular motion. The greater the kinetic energy of a system, the higher the temperature. So we can refine the definition of temperature by saying that it indicates the average kinetic energy of molecules.

Energy Is Transferred in Force Fields

We have found that electric, magnetic, and gravitational force fields surround objects and exert an influence at a distance. We discovered that in the gravitational field, potential energy of position may be translated into energy of motion. Kinetic energy in turn may become potential energy.

We found that if two magnets were suspended near each other, energy could be transferred through the interaction of their fields. When one magnet, hanging from the threads, was set in motion, the other responded and much of the motion was exchanged.

Studies with the electroscope indicated that electrically charged objects also transfer some of their energy to other objects located at a distance. Electrons, which show both electrical and magnetic properties, can transmit energy of motion to other electrons. The intense fields that accompany the buildup of charge concentrations before a lightning discharge strip electrons from the air molecules. If you have ever listened

Hydrochloric Ac
Ammonia Water

to a radio during an electrical storm you may recall that static often accompanies a flash of lightning. This disturbance indicates that not all of the energy associated with the lightning is in the form of light.

Chemical Changes Link Matter and Energy

The chemical changes we have observed, beginning with the formation of crystals in Chapter 4, the cloud of ammonium chloride in our study of gases in Chapter 6, and the chemical reactions we studied in Chapter 11, lead us to believe that these changes occur when molecules of different substances come in contact. We have referred to these contacts as collisions and apply the term *collision theory* for chemical reaction. The collision of moving particles possessed with considerable energy seems to be a very necessary part of chemical change. But mere collision is not enough. The energy of the collisions must be sufficient to cause an interaction to occur; electric charges must be disturbed from their normal condition and the fields of the particles must interact. This means that the rate of many chemical reactions is determined by the temperature of the system. Remember, temperature is an indication of the average energy of motion of the molecules. Increasing the temperature of the substance increases the probability that chemical reactions can occur. Other factors such as concentration and surface area also determine the opportunity for collision. Catalysts affect the reaction conditions so that the energy necessary for interaction is reduced.

Some chemical substances release energy and others absorb energy as they react. A burning candle releases energy and the chemicals in a storage battery absorb energy as the battery is recharged. If we apply the idea of potential energy to chemical substances, we can say that they either gain or lose potential energy as they react. The materials that result from the reaction either contain more or less potential energy than the reacting substances. The gains or losses in potential energy represent adjustments in the relationships of the force fields of the particles in the molecules of the reacting substances.

12–3 THE KINETIC-MOLECULAR THEORY: OUR BEST EXPLANATION

When we *explain* something, we build a bridge between the familiar and the unfamiliar by using mental pictures and theories and by making comparisons. We gather facts during experiments, but facts alone are not science. Data must be interpreted; to do this successfully, we invent concepts or theories. Our remembered experiences, our attitudes, our beliefs, all influence the theories we build. Sometimes concepts prove to be inadequate, as did the caloric fluid theory for heat and the continuous theory for matter. When this happens we either change the concept or discard it entirely and replace it with a more satisfactory one.

A good theory ties together many isolated facts in a simple and reasonable way. It also suggests new experiments to try and helps to stimulate new ideas. Theories are a way to unify knowledge and give it meaning.

At the present time we accept the particle-motion theory as being the best way to explain many of the phenomena we have observed during the long time we have been accumulating facts. This theory is commonly called the *kinetic-molecular theory*, because it uses the idea that matter is made of molecules that are in constant motion. This theory relates matter and energy in a reasonable way and provides a firm basis for interpreting a large body of observations and facts.

The behavior of matter in all its states (solid, liquid, and gas) can be explained most satisfactorily if we assume it to be composed of small particles and that these particles possess kinetic energy. This theory also enables us to make predictions about how matter may react under untried circumstances. Because of its great usefulness, the kinetic-molecular theory of matter and energy is one of the most powerful tools at our command.

12–4 NATURAL TENDENCIES IN THE UNIVERSE

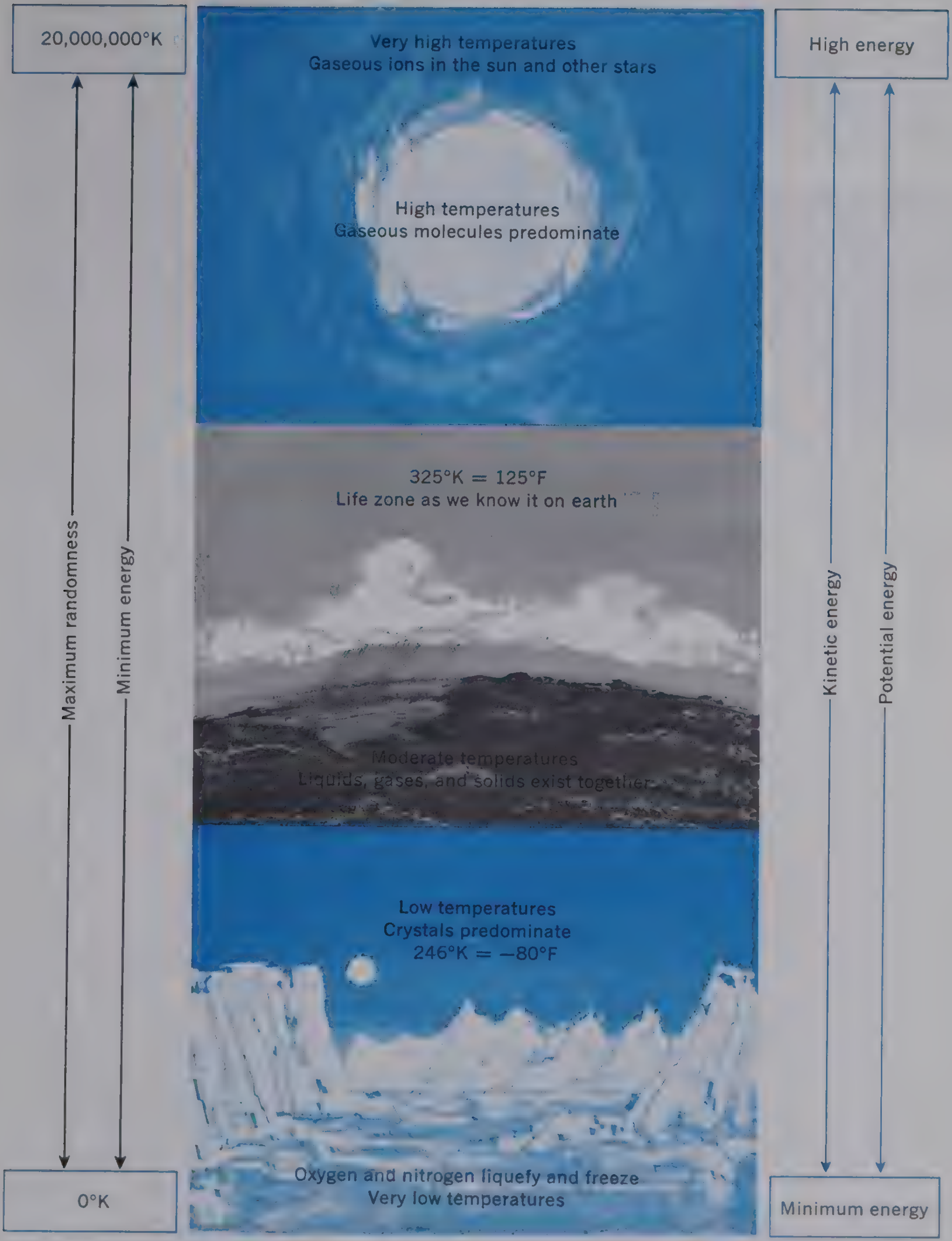
In Chapter 5, as we studied heat, we talked about randomness and order in nature. Now that you have had more experience with energy and the changes it produces, we should return to this concept for further consideration.

You know that water flows downhill. You performed experiments in which you found that heat energy is always transferred from hot objects to colder objects. A pendulum set in motion soon comes to rest. Warm objects, such as the tube containing zinc and copper sulfate, soon cool. Crystals dissolve in water. Gases slowly spread through a tube of air. It has become increasingly evident that many reactions tend to proceed from orderly patterns to less orderly arrangements. The dissolving of crystals is one example. It is also evident that matter tends to go from higher energy states to lower energy states. Most of the other examples cited above are indications of this.

Light patterns from an irregular reflecting surface combine randomness and order in one situation.

The evidence from a great number of investigations has led scientists to conclude that there is a general tendency for materials to move toward the most random distribution of particles and also toward the condition of lowest possible energy. This is true for both physical and chemical changes. Frequently, randomness and lowest energy are conflicting tendencies and both cannot be achieved at the same time. In such cases, equilibrium is reached at intermediate levels of randomness and energy. In crystals, the lower energy states are found when the materials have reached a certain orderly arrangement and for these solids the crystalline form is apparently preferred. If crystals of water (ice) are heated, the water will take up the more random nature of a liquid. This is one definition of melting. It is the temperature at which added energy increases the disorder rather than the kinetic energy of the particles. Conversely, freezing is the act of sacrificing the characteristic chaos of the liquid phase in order to accommodate the tendency to reach the lowest possible energy level.

Life seems to be a process in which the two most evident tendencies of the universe (lowest energy and least order) are violated. In all living things, common, inanimate substances become uniquely organized. They do so by the addition of energy rather than by its loss. True, the energy continually wastes away, and, to maintain the organization of life, it is necessary to restore this energy, either in the form of radiation received by plants or the foods taken in by other living things. A living thing dies when it gives in to these universally dominant trends. When life passes, the ability of the organism to store energy in its highly organized structures is lost and both the organization and the energy begin to decrease.


20,000,000°K
High energy
Very high temperatures
Gaseous ions in the sun and other stars
High temperatures
Gaseous molecules predominate
325°K = 125°F
Life zone as we know it on earth
Maximum randomness
Minimum energy
Kinetic energy
Potential energy
Moderate temperatures
Liquids, gases, and solids exist together
Low temperatures
Crystals predominate
246°K = −80°F
Oxygen and nitrogen liquefy and freeze
Very low temperatures
0°K
Minimum energy

Man creates order in physical materials and in social structures.

Notice, in passing, that these same principles apply to social structures—organizations of living things. The natural tendency of the units (individuals) of a social organization, considered only as physical entities, is toward randomness. But there are other forces that overcome these tendencies; they result in the production of colonies of plants or animals with a high level of organization, or the social structures of human civilization, such as football teams, marching bands, cities, and nations. Like the living things of which they are composed, these social structures are contrary to the nature of the inanimate material and, when energy is no longer supplied, they revert to senseless scattered rubble.

12–5 THE SUN: OUR SOURCE OF ENERGY

We have seen that the rate of energy transfer in a system is greatest when the energy difference between the parts of the system is also great. A very hot object gives up energy at a more rapid rate than an object at a moderate temperature. If it were not for the tremendous energy source we call the sun, the earth would probably have come to an intermediate energy level long ago. If this had happened, no plants could grow to furnish fuels for animals or machines, and no water would evaporate to fall again upon the earth. Under such conditions the earth would be a lifeless, dormant mass, responding only to the convulsions of a hot interior. When the earth had spent its energy reserve, no observable change of any kind would then be expected. We are totally dependent on the continuing transfer of energy from the sun to the earth.

What is the source of energy from the sun, the prime mover of our solar system? Scientists now believe energy results from the combining or the dividing of particles of matter that are even smaller than the molecules and atoms we have discussed. In addition to being able to provide the enormous energy we associate with our sun, the reactions of these tiny particles provide the energy we associate with the glow of a luminous watch dial. Energy of this kind is called *nuclear energy*.

Man's increasing ability to exploit nuclear energy has opened up a new era of possibilities for advancing technology as well as making greater responsibilities for society necessary. Man is now able to use the ultimate source of the energy of the universe and must exercise great judgment in deciding just how it will be used. The power to create is no different from the power to destroy; it is only a matter of the *direction* in which this power is expanded.

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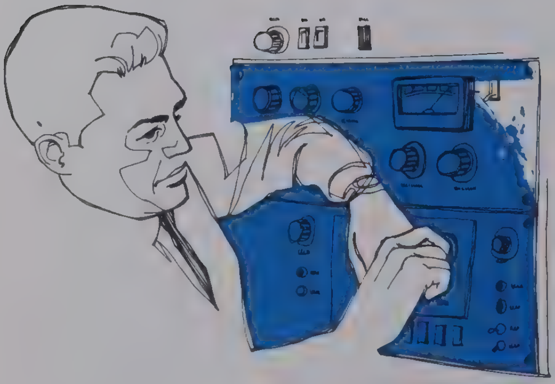
13

Looking Ahead to Future Study of Matter and Energy

13–1 Scientific Knowledge Is Tentative
13–2 Science and Philosophy
13–3 Your Future in Science
13–4 Responsibility and Challenge



The building shown on the opposite page is two miles long and houses a linear accelerator used to study the nuclei of atoms.

13–1 SCIENTIFIC KNOWLEDGE IS TENTATIVE

The search for fundamental understanding of our environment brings us to an end point. We have built a theory for matter and a theory for energy by which we have explained our observations. We are more comfortable about this knowledge because we have gathered it ourselves. But our knowledge is not a complete package. Our investigations were directed at obvious, simpler things which are within our ability to observe without elaborate instruments. Our measurements were not always precise or very quantitative. Is there another "world," either too small or too distant, where matter does not follow the general statements we have composed? If we could measure density, force, or temperature more carefully, would we then have to alter our statements? Are there relatively common and simple systems we did not study that might be different enough to render our models useless in explaining their structure or behavior? We have learned enough about the nature of science to realize that any extension of our view of matter and energy must be done very cautiously.

The work we have done is important, but it is not complete. Our theories are of value as a means of relating many observations and in predicting what we might expect when we go on to investigate other systems. Our present knowledge of the method and content of science is a good point of departure for future study. All we have learned should be regarded as tentative. It must be subject to constant review and refinement as new evidence and new ideas appear.

We have taken a short tour into the area of scientific knowledge. This may have seemed to be a strange land, but perhaps you have felt a challenge to explore more deeply its beauty and its customs.

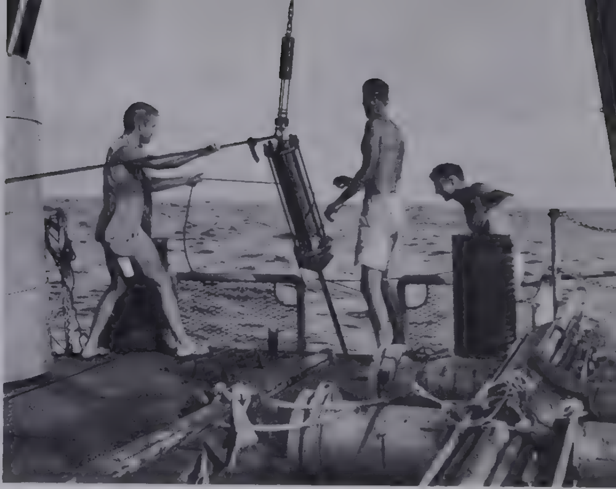
What can we say about future explorations? To date we have a description of a pattern based on what we have observed. If the pattern seems to be valid as we attempt to encompass more and more systems, then we may try to weave in more threads to help us understand not only "what" but "why." A theory about why matter behaves as it does is more useful in predicting new situations than a simple description or an orderly grouping of facts. The development of theory is an adventure of the mind; it exercises the highest capabilities of man. Through this door we may pass to new levels of understanding and enjoyment. And in doing so we may develop increased faith in the intricate order of the entire universe. If we can understand this order and its unity, perhaps it will have simplicity. Simplicity made evident by unity is the keynote of scientific explanation.

13–2 SCIENCE AND PHILOSOPHY

The further we go in trying to explain the "why" of nature, the more closely we approach the field of philosophy. Science is centered in the "what" and "how" of nature. However, science makes excursions into the "why" each time a theory is formulated. These excursions are really not very deep. They help us to link our many isolated bits of knowledge, but as yet they have not brought us very close to ultimate answers. But human beings are optimistic creatures.

We may never establish a complete explanation for such basic concepts as the nature of the electric charge and the behavior of charged particles. These things have always existed. The answer, "because that is the way they are," to the question of "why" is considered unacceptable and unscientific. It is more desirable to leave the question open. We should continue to strive for the greatest possible understanding. We should not hestitate to ask questions, even if we are not sure we can ever find an answer. We can start along any path we choose, even though we have no assurance that we shall reach its end. We may help to establish base camp on the route which will make it possible for others to reach the summit. Like so many explorers, we may see from the top of the hill ahead a view of a higher and more difficult crest beyond. Thus, we may find both a deep satisfaction from our achievements and the challenge of a higher goal.

What are the paths we might follow as we explore what is known and reach toward the unknown? Like any exploration, it would largely depend on where we wanted to go. To those interested in the oceans of the world, the logical course would be to the water's edge. The mountaineer goes in another direction with different tools to aid him. Both must eventually realize, however, that parts of the region they study are affected by other regions. Mountains are carved and eroded by water and ice; water, which came from the oceans and will someday return to them, is rich with dissolved and suspended matter wrested from the mountains. Within a



The oceans are a scientific frontier that we are just beginning to study. Here scientists lower a temperature probe into the ocean to study the amount of heat flowing through the crust of the earth into ocean waters.

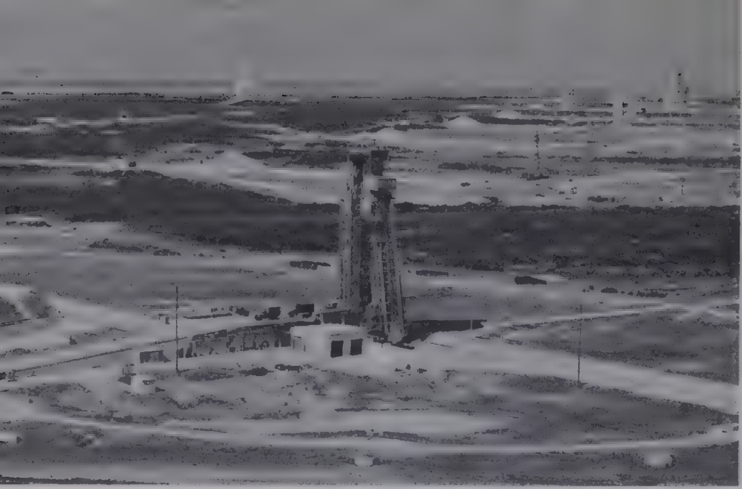
system such as our earth, no set of circumstances is entirely unrelated to other seemingly separate circumstances. Everything we experience can be understood most fully only when considered in the context of everything else. As we continue to study we will be torn between the necessity to specialize in a relatively narrow field, and the importance of looking at this narrow field in its relationship to the larger body of knowledge.

13–3 YOUR FUTURE IN SCIENCE

The need to specialize as we advance in science studies has led us to classify the different areas. First, we divide scientific knowledge into the biological and the physical sciences. Living things fall into the category of biological science, or life science. The physical sciences are concerned more with matter and energy, the behavior of which appears to be more constant and predictable. This classification is subject to the same weaknesses as any rigid system of classification. There is great overlapping between the two divisions. For example, the action of matter and energy in living systems can be predicted by using the same physical laws that are used to predict the action of matter and energy outside of living systems.

The living systems affect the nonliving systems and vice versa. We know that mountains would not be the same if it were not for the plants and animals that populate them; nor would living things exhibit their present characteristics except for the physical environment surrounding them.

The extent to which you decide to go in your study of science will determine how you select your future studies. We have looked at science in general. Had we held primary concern for living things, we would



One of many facilities at Cape Kennedy. Space exploration is just one of many scientific fields opening new areas of study for future scientists.

have called our subject biology. Following this same topic, at a higher level of specialization we might concentrate on the plant kingdom in a study of botany, or on the animal kingdom in a study of zoology. From each of these we could continue to branch out to more specific areas where we might consider increasingly narrower groups of organisms. We should establish a broad foundation upon which to build before we become too highly specialized.

Those persons more interested in the nonliving division of science will find a similar network of subjects. The main concern in a study of physics is energy; chemistry is concerned with matter and its changes; but, again, there is much intertwining. In fact, much of the research being done at the present time in the study of living things deals with the molecules that compose living tissues. This has become the joint concern of molecular biologists, biochemists, and biophysicists. The more we learn about living systems, the more we are coming to realize that a knowledge of physics and chemistry is an essential foundation for understanding the interactions that occur in plants and animals. Man's response as he makes excursions into the new environment of outer space involves both his body chemistry and the new forces to which he will be subjected. Because of the intersection of physics, chemistry, and biology this new area of specialization has been termed "space medicine."

There are two significant aspects of increased scientific specialization that lead to problems which must be resolved. First, as a person increases the depth of his study, he must necessarily reduce its breadth. Thus, he will have less opportunity to see his work in the context of all knowledge, or to understand the work of others. Second, the new vocabulary of a subject, with the specific meanings words may have in application to a narrow field, makes it increasingly more difficult for people to communicate with one another. The language of a scientist as he discusses his area of specialization is almost as strange to the average person as if it were foreign and written with a different alphabet. The tendencies produced by these aspects of scientific effort need to be counterbalanced by an effort to perceive the threads of theory that can unify the diverse work of all scientists. This is essential if physicists, chemists, and biologists are to work effectively by bringing their special knowledge to bear on the problems of an area like space medicine.

13–4 RESPONSIBILITY AND CHALLENGE

Many of you have found from this study that the pursuit of science can result in a most satisfactory and rewarding life. If you are willing to devote the time and effort to adequate preparation, you may play an important part in the advance of knowledge. Some of you may choose to become scientists. Others may help to fill the growing demand for broadly educated people who can interpret science to the layman. These are the teachers, writers, and speakers of the scientific and technological fields — those who pass science on to future generations. All those who work in science and related fields are spinners of dreams who can become powerful instruments for good as they weave and reweave the intricate patterns in the fabric of universal knowledge.

Whatever may be the work to which our lives are dedicated, certain responsibilities are clear. We must seek the solutions to all kinds of problems, both scientific and social. No matter what part we are destined to play in the future of mankind, we must keep informed of the kinds of developments that are being made in all areas of learning. We must be aware of the particular role of science in both the advance of knowledge and in shaping the events of individual lives. We must strive to understand the nature of science itself, both as an adventure of the mind and as a significant force in the advance of civilization.

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Glossary

acetic acid Acid which contains hydrogen, oxygen, and carbon and is present in vinegar. In a water solution, it conducts an electric current.

alum Aluminum potassium sulfate and similar compounds that contain water as part of the crystal.

ammonium chloride A compound often called by the common name sal ammoniac; absorbs heat during the process of dissolving in water.

amplitude The extent of a vibratory motion, as in a pendulum. It is the maximum distance the bob moves out from the point of lowest potential energy.

anhydrous Describes a compound which contains no water.

anneal To toughen metal by heating to a high temperature.

aspirator A suction pump used for causing fluids to move by changing the pressure on them.

barium chloride A compound which is soluble in water and is composed of one part barium and two parts chlorine.

battery A group of cells connected together for producing current.

bob A weight which hangs from the end of a line, as the bob on a pendulum.

butyl carbitol A carbon compound of

diethylene glycol and monobutyl ether; its density is 0.95 g/ml.

calorie A derived unit for measuring heat. The amount of heat absorbed by 1 g of water in raising its temperature 1°C; or the amount of heat released as 1 g of water cools by 1°C.

Calorie A unit for measuring heat equal to 1000 small calories. Also known as the large Calorie, or kilocalorie.

calorimeter A container used to measure quantities of heat.

catalyst A substance which changes the rate of a chemical reaction.

cell In electricity, a device which can convert chemical energy into electrical energy.

Celsius scale A thermometer scale which divides the interval between the freezing and boiling points of water into 100 units or degrees. Also called the centigrade scale.

chemical energy The change in energy that accompanies a change in the composition of one or more substances.

chemical reaction A change in which one or more new substances are formed.

compound A complex substance formed of two or more elements.

concentrated solution A solution in which solute is present in a large amount compared to a dilute solution; not a precise term.

conduction A form of heat and electrical transfer that does not involve a flow of the substance.

conductor A substance which transfers heat or electricity readily.

controlled experiment A two-part experiment with only one factor that differs; also, a carefully regulated experiment.

convection A method of heat transfer in fluids by current flow.

copper chloride A blue-green substance composed of copper and chlorine.

copper nitrate A blue-green substance composed of three elements—copper, nitrogen, and oxygen.

crystal The form of most solids; particles are arranged in a definite and characteristic pattern.

deduction A reasoning process in which past occurrences are used as the basis for prediction of future occurrences.

density A property of matter determined by dividing the mass by the volume; it is a derived value.

diffusion The random scattering and intermingling of substances due to particle motion.

dilute solution A solution in which solute is present in a relatively small amount compared to a concentrated solution; not a precise term.

elastic Tendency of matter to return to its original condition after distortion.

element A simple substance which cannot be separated further.

electrolysis The separation of a chemical compound into simpler substances by action of an electric current.

electrolyte A liquid or solution which conducts an electric current.

electron A fundamental particle found in all atoms.

energy The ability to do work; the action of a force through a distance.

enzyme Substances found in living organisms that act as catalysts.

equipotential line A line in a force field along which the strength of the field is constant.

etching The process of using a corrosive substance to treat the surface of a metal.

extrapolation The process of predicting data beyond the observed range.

Fahrenheit scale A thermometer scale on which (at standard conditions) the freezing point of pure water is 32° and the boiling point of pure water is 212°.

ferrous sulfide A compound of iron and sulfur.

field A region around a source of magnetic, electrical, or gravitational force where the force can be detected.

field lines Lines which identify the direction of the force in magnetic, electric, and gravitational fields.

filtrate The solution portion of a liquid mixture; it can pass through a filter paper.

friction Resistance to motion of one object rubbing or rolling over a second object.

fulcrum Pivot or point of support for a lever; axis of rotation.

germanium An element with properties similar to carbon and silicon. Useful in making semiconductors.

goniometer An instrument used to measure angles between faces of a crystal.

heat sensor A nerve ending which is especially sensitive to temperature changes.

hydrometer A floating instrument used to determine the density of liquids relative to water.

hypo A common name for sodium thiosulfate, a substance used in photographic processes. Also, a prefix meaning *under* or *lower*.

hypothesis A tentative answer by which one explains observations and which guides plans for further investigations.

induction A reasoning process which uses specific observations to identify a pattern or rule which can be used as a general basis for deduction.

insulator Material which does not transfer heat or electricity readily.

ion A particle which has acquired an electric charge.

Kelvin scale A temperature scale on which the zero point is about −273°C; sometimes called the Absolute Scale.

kilogram-force Force of attraction acting between the earth and a one-kilogram mass. In physics, it is called a "newton."

kindling temperature The lowest temperature at which a substance begins to burn.

kinetic energy Energy of motion.

krypton An element of the group known as the rare gases.

the law of universal gravitation A generalization proposed by Newton which states that every object in the universe exerts a gravitational force upon every other object.

machine Any device which modifies, transforms, or transmits energy.

magnesium ribbon A thin strip of the metallic element, magnesium.

mass A term used to denote the quantity of matter in an object.

metric system A decimal system of measurement based on the meter as the unit of length and used in all scientific work.

mossy zinc A form of the metallic element, zinc, which contains carbon as an impurity.

nickel sulfate A green crystalline substance which contains the elements nickel, sulfur, and oxygen.

nonelectrolyte A liquid or solution which does not conduct an electric current. For example, alcohol and sugar in water.

nuclear energy Energy changes resulting from changes in the nucleus of atoms.

octahedron A solid with eight faces.

opaque Unable to transmit visible light.

orderliness Regularity or pattern in the arrangement of particles or objects.

osmosis Process of diffusion in which water passes through a semipermeable membrane from an area of greater concentration of water to an area of lesser concentration.

ozone A form of oxygen consisting of three atoms per molecule instead of the usual two atoms per molecule. Produced by passing an electric current through oxygen.

physical change A change in which the appearance of matter is altered by the composition of the substance. For example, dissolving sugar in water is a change which alters the appearance of the sugar, but does not permanently change it. The sugar can be recovered from the water by evaporating the water.

plasma The ionized, gaseous condition characteristic of matter at very high temperatures; sometimes called the "fourth state of matter."

plastic A property of matter which describes its ability to be molded or distorted; nonelastic.

poles The two points on a magnet which attract iron filings with the greatest force.

potassium chlorate A white substance composed of potassium, chlorine, and oxygen.

potassium ferricyanide A red-brown substance composed of potassium, iron, carbon, and nitrogen.

potential energy Energy of position. For example, pendulum bob at the highest point in the arc.

precipitate A relatively insoluble solid produced by mixing two solutions. For example, when magnesium sulfate and barium chloride solutions are mixed, a white solid forms (barium sulfate).

radiation A form of heat transfer by wave motion.

radiometer A device for demonstrating that radiation moves matter and is therefore a form of energy. Consists of a partially evacuated glass bulb in which a set of four vanes are mounted so that they are free to turn.

residue The solid caught on the filter paper during the process of filtering.

Rochelle salt White salt composed of potassium, sodium, carbon, hydrogen, and oxygen suitable for easy crystal growing. Its chemical name is potassium sodium tartrate.

salol (phenyl salicylate) A low melting-point solid which rapidly forms crystals.

salt One of a class of chemical compounds formed when the hydrogen in an acid is replaced by a metal. Water solutions of salts are good electrolytes.

scalar A quantity which can be represented by a point on a scale; having magnitude only, not direction. For example, speed.

seed crystal A tiny crystal which when suspended in a solution of proper concentration may serve as the base on which a large crystal can grow.

silicon An element present in the substances that comprise glass and sand; useful in making semiconductors.

silver nitrate A white crystalline compound composed of silver, nitrogen, and oxygen.

sodium sulfate White crystalline substance composed of sodium, sulfur, and oxygen.

solute The dissolved substance in a solution.

solvent The dissolving portion of a solution.

specific gravity A ratio which compares the density of a substance (usually a solid or a liquid) and the density of water.

speed Rate of travel; distance moved divided by time; a scalar quantity.

stalactite A deposit of calcium carbonate which hangs from the ceiling in caverns.

stalagmite A deposit of calcium carbonate which forms from the floor in caverns.

static electricity Stationary charges of electricity produced by rubbing objects.

sulfur, amorphous That form of the element sulfur which does not display a clearly crystalline pattern.

sulfur, monoclinic One of the crystalline forms of the element sulfur; crystals are slender and needlelike. Also called prismatic sulfur.

sulfur, prismatic *See* sulfur, monoclinic.

sulfur, rhombic The most stable form of sulfur in which the crystals are diamond-shaped.

taxonomy The science of classification.

tetrabromoethane A liquid carbon compound containing bromine with a density of 2.96 g/ml.

tetrahedron A solid with four faces.

torque Force which produces or tends to produce rotation about a point.

transparent Able to transmit visible light.

Tyndall effect Scattering of light by small particles. Demonstrated, for example, in the beam of a car headlight on a foggy night.

vapor The gaseous form of a substance which is a liquid at room temperature and pressure.

vector A quantity that has both magnitude and direction.

weight A measure of gravitational force upon an object.

work Transference of energy; measured in kgf-m.

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